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ABSTRACT

This book contains 112 classroom demonstrations intended to be used with any introductory chemistry program. These demonstrations were selected in an effort to provide simple, safe, effective and enjoyable experiences for the class. In addition, they are intended to be used to introduce many of the major concepts in chemistry. The demonstrations involve color changes, gas evolution, precipitate formation, smoke, fire, and other obvious or dramatic chemical changes. The guide is organized into 11 major sections including: (1) introducing chemistry; (2) physical changes; (3) reactions involving gases; (4) reactions of some elements; (5) transition metals and complex ions; (6) chemical bonding; (7) energy changes; (8) solutions and solubility; (9) kinetics and equilibrium; (10) acids and bases; and (11) oxidation and reduction reactions. Appendices include an equipment and reagent list and detailed safety and disposal instructions. (TW)

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Chemical Demonstrations

A Sourcebook for Teachers

Volume 2

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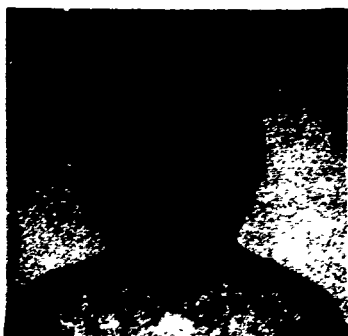
science educational development programs and as a chemistry consultant to various National Science Foundation Institutes. He has presented many seminars and published a number of books on methods and aspects of teaching chemistry and science. He has also conducted numerous workshops on chemical demonstrations throughout the country. He served as coordinator of the Institute for Chemical Education program at the University of California—Berkeley in the summer of 1986. Dr. Summerlin has held office or had major committee assignments in the American Chemical Society, the National Science Teachers Association, the American Association for the Advancement of Science, the Florida Association of Science Teachers, and the Association for the Education of Teachers of Science. He was the recipient of the American Chemical Society Outstanding Chemistry Teacher Award (1967), the James B. Conant Award (1969), and the Du Pont Award (1970). He received the 1985 Ingalls Award, which is given to the Outstanding Teacher at the University of Alabama at Birmingham, and was a national recipient of the 1986 Chemical Manufacturers Association Catalyst Award.



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activities and demonstrations for science-shy students, including high-school dropouts. She has helped conduct advanced-placement chemistry workshops at several sites including the Hill School in Pennsylvania and has conducted other workshops in chemistry and chemical demonstrations locally and throughout the country. She has presented papers and original chemical demonstrations at several regional and national conventions including CHEM Ed 85 at Monclair State College and CHEM Ed 86 at Montana State University. Ealy served as field editor for *Chemical Demonstrations: A Sourcebook for Teachers* and also served on the American Re-Write Committee and Editorial Board for *ALCHEM*, the Alberta Chemistry Project. Ealy is active in several professional organizations, including the American Chemical Society. She did volunteer work with Subcommittee II of the ACS to select and edit questions for the Advanced Chemistry Examination.

A Word of Caution

Demonstrations can be fun, and they add excitement to the teaching of chemistry. However, this volume is intended for use by professional chemists and chemistry teachers. Even the simplest demonstration is potentially dangerous when performed by someone lacking the manipulative skills and knowledge of chemistry necessary to understand the reactions involved. Every precaution must be taken to ensure the safety of the demonstrator and the students. You should follow the directions given for each demonstration and not exceed the recommended amounts of chemicals.

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Preface

The success of the first volume of *Chemical Demonstrations: A Sourcebook for Teachers* far exceeded any of our expectations. We are certainly pleased with this response, because it implies that we are perhaps on target with the selected demonstrations. However, we are much more pleased because this success indicates that chemistry teachers are putting forth an extra effort to gather material to improve their teaching of chemistry. We are impressed that so many teachers are receptive to new ideas and techniques and are willing to share with us some of their thoughts on improving instruction in chemistry. This second volume of *Chemical Demonstrations: A Sourcebook for Teachers* is an outgrowth of this positive attitude and these ideas and suggestions.

This second volume contains more than 100 demonstrations appropriate for an introductory chemistry program. Following the philosophy and format of the first volume, these demonstrations are simple, safe, effective, and enjoyable. The demonstrations are presented in a simple format for quick reference and use. We have purposely left out lengthy discussion of chemical principles, extensive references, and detailed reactions. You can find these in any general chemistry textbook, and you are encouraged to do so. We feel that it is most useful for the teacher to be told what the demonstration does, how to do it, what the reactions are, and how to prepare materials and solutions. We also offer a few teaching tips suitable for use with the demonstrations. We encourage you to use this book as a reference manual—add your own notes, insert additional demonstrations, and make changes and modifications that allow the demonstrations to become a personal part of your chemical repertoire. We also include several demonstrations of historical interest—including the oxidation and reduction reactions of indigo, the oldest known coloring matter; Ira Remsen's experience with copper and nitric acid; and even the oldest recorded demonstration! We hope that you enjoy using these demonstrations with your students.

Because many demonstrations illustrate several important chemical principles, and applications of those principles, we have provided a listing of demonstrations and chemical topics as Appendix 1. This listing will help you to quickly find a demonstration to fit a specific topic.

You should always take necessary precautions to ensure the safety of yourself and your students. We have purposely excluded several very dramatic demonstrations because we do not consider them to be safe and appropriate for use in the introductory chemistry program. If special precautions are needed, these precautions are presented with the appropriate demonstrations. We have excluded, whenever possible, hazardous chemicals. Our directions regarding the use and disposal of hazardous substances comply with the standard chemical disposal procedures as recommended in *Prudent Practices for Disposal of Chemicals from Laboratories* (published by the National Academy Press, 1983) and the U.S. Consumer Product Safety Commission guidelines as listed in *School Science Laboratories: A Guide to Some Hazardous Substances* (1984; see Appendix 4).

Performing demonstrations in chemistry can be an exciting and enjoyable experience for the teacher and an enlightening and intriguing experience for the student. To obtain the maximum effect from your demonstrations, we offer the following suggestions:

1. Always rehearse your demonstrations before doing them before a class. Sometimes demonstrations work, and sometimes they do not. You should work out all of the "bugs"—check the purity of your chemicals, adjust the amounts of the chemicals, and try variations. When you finally get a demonstration to work to your satisfaction, carefully note everything you did to make it "go".
2. Always try to actively involve students in your demonstrations. Have several students assist by mixing solutions, reading thermometers, recording data, and so on.
3. Demonstrations should be fun. They should catch and hold the student's interest.
4. Demonstrations should not be thought of as a replacement for a laboratory experience. Nothing substitutes for hands-on experiences provided by the laboratory. Demonstrations offer an effective way to extend the laboratory experience and help students develop skills in observing, predicting, recording, and analyzing data.

Most demonstrations in chemistry have been around for a long time—some of them literally for hundreds of years. These demonstrations have undergone many modifications and changes—including ours—that make it almost impossible to find and credit original sources. Therefore, we simply express our appreciation to all of the chemists, past and present, who have developed and used chemical demonstrations to make chemistry more meaningful and more enjoyable. We also express our appreciation to the hundreds of teachers in our workshops and chemical demonstration programs who have used our demonstrations and offered suggestions for improvement.

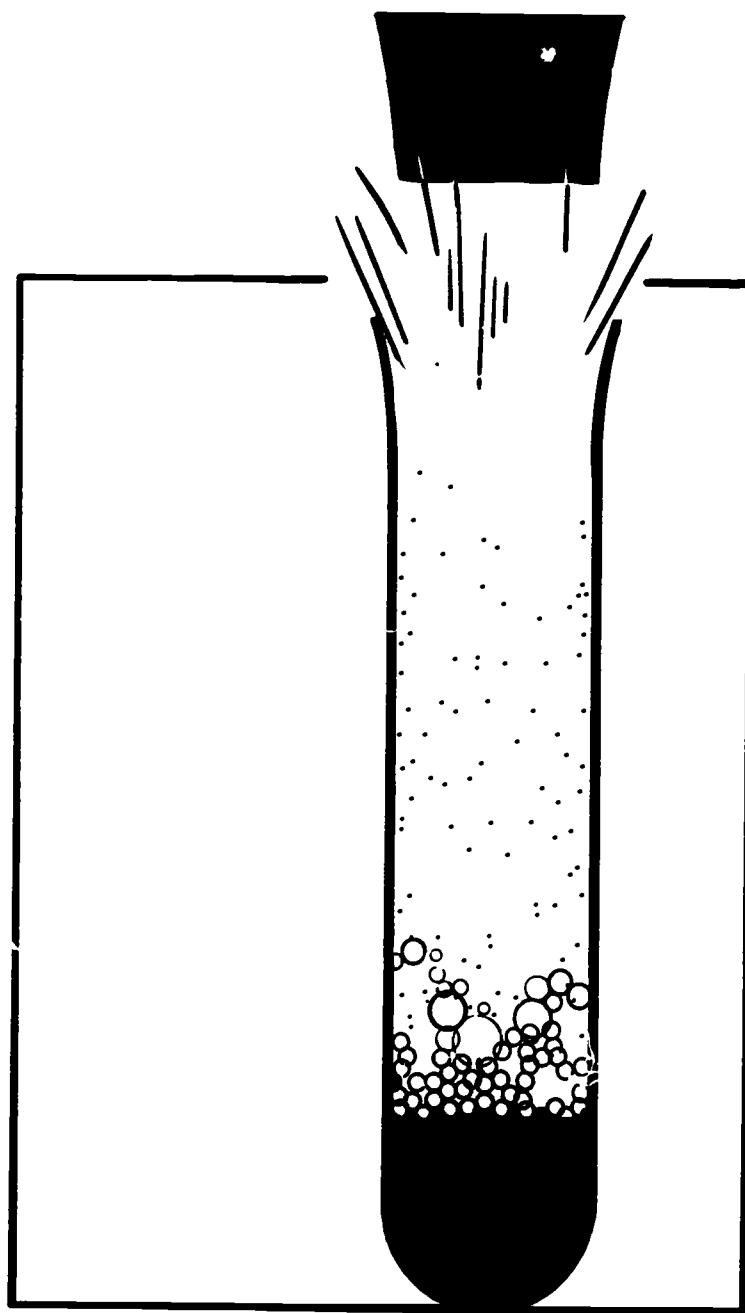
Special thanks are due to Jim Ealy, coauthor of the first volume and special consultant for this volume, who carefully reviewed the manuscript and made valuable suggestions for improvement; our editor at the American Chemical Society (ACS); and Bob Johnson, Acquisitions Editor for ACS.

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Introducing Chemistry



1. The First Demonstration: Proof That Air Is a Substance

This lecture-demonstration is the first recorded in history. The demonstration was performed in 440 B.C. by the Greek physician Empedocles to confirm his belief that air is a “substance”. With minor changes, this demonstration can be repeated.

Procedure

1. Fill a large container almost full of water.
2. Add a few drops of food coloring to make the level more visible.
3. Place your finger over the opening in the stem of a funnel and place the large open end of the funnel in the water.
4. Notice that the water does not enter the funnel.
5. Remove your finger from the opening. Notice, as Empedocles observed, that “the water rushes into the large opening as the air rushes out of the small”. Because air prevented the water from entering, air must, indeed, be a substance.

Teaching Tips

NOTES

1. On the basis of this observation, Empedocles developed a “water clock”. This clock was a conical container with a small hole in the top and a larger hole in the bottom. The cone sank in a certain amount of time, so it could be used as a crude time piece.
2. Ironically, this first demonstration can be easily documented and attributed to Empedocles more than 2300 years ago, although the origins of more recent demonstrations are much more difficult to trace.
3. By modern standards, this demonstration seems very simple and obvious. However, it was not the custom of early “scientists” to do experiments or demonstrations. This simple demonstration by Empedocles caused other Greek physicians to begin thinking about the role of air in breathing and other processes.

QUESTIONS FOR STUDENTS

1. Why do you think such a simple demonstration was important in the year 440 B.C.?
2. Would you give the same explanation as Empedocles?
3. Find out more about Empedocles, the Greek physician.
4. What is a *demonstration*?

2. Ira Remsen's Investigation of Nitric Acid

In this classic demonstration, the experience Ira Remsen had with nitric acid can be reenacted. Ira Remsen (1846–1927) was an influential chemist in America. He founded the chemistry department at Johns Hopkins University and initiated the first center for chemical research in this country. In this demonstration, he describes his experience with nitric acid.

Procedure

It is suggested that you read his account as you perform the demonstration.

While reading a textbook on chemistry, I came upon the statement "nitric acid acts upon copper". I was getting tired of reading such absurd stuff and I determined to see what this meant. Copper was more or less familiar to me, for copper cents were then in use. I had seen a bottle marked "nitric acid" on a table in the doctor's office where I was then "doing time"! I did not know its peculiarities but I was getting on and likely to learn. The spirit of adventure was upon me. Having nitric acid and copper, I had only to learn what the words "act upon" meant. Then, the statement, "nitric acid acts upon copper" would be something more than mere words.

All was still. In the interest of knowledge I was even willing to sacrifice one of the few copper cents then in my possession. I put one of them on the table; opened the bottle marked "nitric acid"; poured some of the liquid on the copper; and prepared to make an observation.

1. Place the penny in a 500-mL flask.
2. Carefully add 5 mL of concentrated nitric acid.
3. Place several damp, folded, paper towels over the top of the flask to catch the gas evolved.

But what was this wonderful thing which I beheld? The cent was already changed, and it was no small change either. A greenish blue liquid foamed and fumed over the cent and over the table. The air in the neighborhood of the performance became dark red. A great colored cloud arose. This was disagreeable and suffocating—how should I stop this? I tried to get rid of the objectionable mess by picking it up and throwing it out of the window, which I had meanwhile opened. I learned another fact—nitric acid not only acts upon copper but it acts upon fingers. The pain led to another unpremeditated experiment. I drew my fingers across my trousers and another fact was discovered. Nitric acid also acts upon trousers.

4. Display to the class a large piece of cloth on which a dropper full of nitric acid had previously been squirted.
5. Add water to the nitric acid to stop the reaction.
6. Rinse the beaker and remove the penny (now much smaller!).

Taking everything into consideration, that was the most impressive experiment, and, relatively, probably the most costly experiment I have ever performed. I tell of it even now with interest. It was a revelation to me. It resulted in a desire on my part to learn more about that remarkable kind of action. Plainly the only way to learn about it was to see its results, to experiment, to work in the laboratory.

Reaction

Nitric acid reacts with copper to produce the brown gas nitrogen dioxide.



Teaching Tips

NOTES

1. This wonderful excerpt is taken from Getman, Frederick H. *J. Chem. Educ.* 1940, 9-10.
2. This demonstration is an excellent way to begin a course in chemistry. Have students find out more about Ira Remsen.
3. This demonstration must be done in a well-ventilated area or in the hood. NO_2 is a toxic gas.
4. Considering when Ira Remsen did this "experiment", have students comment on his investigation.

QUESTIONS FOR STUDENTS

1. Find out more about Ira Remsen.
2. What do you observe when copper is placed in nitric acid?
3. How does a scientist find out new things?

3. Burning Water

A "new" 500-mL Erlenmeyer flask is taken from a box and filled with water from a faucet. The top of the water is ignited, and a flame is produced. The "water" continues to burn for several minutes.

Procedure

1. Prior to performing this demonstration, place a squirt of cigarette lighter fluid in the flask, swirl the flask to distribute the fluid evenly so that the flask appears to be empty, and replace the flask in the original carton.
2. When performing the demonstration, produce the flask from its "original" container as if it were a new flask.
3. Fill the flask to the top with water from a faucet.
4. For a more dramatic effect, add a pinch of salt from a new box or sodium bicarbonate from a new box (let students see you open the new container).
5. Light the top of the flask. The lighter fluid will have floated to the top of the flask, unobserved by the students.
6. Ask for observations (flame, smoky product, kerosene-like odor, etc.).

Reaction

1. The real reaction is that of the students as they try to figure out what is going on. With a tongue-in-cheek presentation, the teacher can produce all sorts of reactions. Some students will believe that the water, salt, or baking soda is burning. Others will rely on their own observations and be skeptical; they know that another ingredient must be involved.
2. The chemical reaction is the combustion of cigarette lighter fluid, a mixture of light hydrocarbons that are "lighter" (less dense) than water. Lighter fluid is naphtha, a mixture of hexanes.



Material

Fluid for cigarette lighters. Do not use charcoal lighter fluid.

Teaching Tips

NOTES

1. If too much fluid is used, students will be able to see an "oily" layer, and the effect is lost.
2. Practice to get just the right amount. With larger flasks, more time is required for the fluid to reach the top.
3. If students assume that the "burning" is a result of some reaction with the salt, or baking soda, write their proposed reaction on the board.
4. This demonstration is FUN! Develop your own story line to go along with this demonstration.
5. Naphtha is distilled from petroleum at 70-90 °C.

QUESTIONS FOR STUDENTS

1. This really special water appears to be burning. Write down all your observations.
2. Using your observations, suggest an explanation for "burning" water.
3. Suggest a way that, using only H_2O , a burning reaction can take place. (Electrolyze the water to hydrogen and oxygen and then react these two substances to reproduce water!)

4. The Copper Cycle

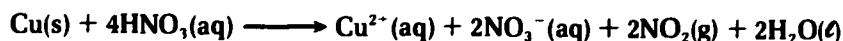
Copper undergoes a series of reactions producing distinctive and colorful products.

Procedure

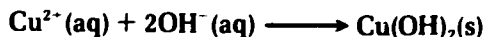
1. Preheat a hot plate or prepare a Bunsen burner for step 7.
2. Place about 1.5 g of fine copper wire or wool in a medium-sized beaker.
3. The nitric acid, HNO_3 , concentration is 6 M: See Appendix 2.
4. The hydrochloric acid, HCl , concentration is 6 M: See Appendix 2.
5. When the first reaction is complete, slowly add 20 mL of 6 M NaOH .
6. Observe a dark blue precipitate and evolution of heat.
7. Place the beaker over the heat until the precipitate is converted to black copper oxide.
8. Decant some of the solution.
9. Copper has been in wide use for 5000 years.
10. See Demonstration 39.
11. Observe the copper forming on the surface of the aluminum.

Reactions

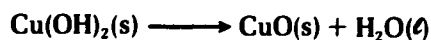
1. Copper is oxidized to copper(II), which is blue in this solution.



2. Copper ions are precipitated as blue insoluble copper hydroxide.



3. Copper hydroxide decomposes to black copper oxide and water.



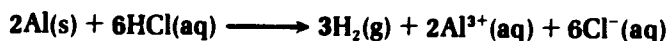
4. Copper oxide reacts with hydrochloric acid to form green copper chloride solution and water.



5. Solid copper is produced in the reaction of copper chloride solution with aluminum.



6. The aluminum also reacts with the acid to form hydrogen gas.



Materials and Solutions

1. Solid copper wool or fine wire.
2. Solid aluminum sheet cut into strips or 18-gauge wire.
3. The nitric acid, HNO_3 , concentration is 6 M: See Appendix 2.
4. The hydrochloric acid, HCl , concentration is 6 M: See Appendix 2.
5. The sodium hydroxide, NaOH , concentration is 6 M: Add 120 g to water to make 500 mL of solution.

Teaching Tips

NOTES

1. Demonstration 2, "Ira Remsen's Investigation of Nitric Acid", could be a dramatic way to introduce the idea of reactions of copper.
2. Copper wool and wire are shiny copper but have enough surface area to react quickly.
3. The nitrogen dioxide gas in the first reaction is toxic. When students have seen the gas, the beaker should be placed in the hood for 1-2 min to allow the gas to be blown out of the room.
4. Each reaction allows time for a student to write the equation on the board. Students enjoy identifying reactants, predicting products, and trying to figure out formulas. Beginning students may prefer to write molecular formulas as $\text{Cu}(\text{NO}_3)_2(\text{aq})$ instead of ionic formulas.
5. You might also want to discuss the neutralization reaction in procedure 5.
6. Copper can be seen to have all reacted in the final step if the solution is colorless.
7. When each step is performed qualitatively, accompanied by discussion, this copper cycle can be completed easily in one class period.
8. If you have taken the mass of the copper at the beginning, you can dry the final product and reweigh it to allow students to calculate percent recovery.
9. Copper has been in wide use for 5000 years.

QUESTIONS FOR STUDENTS

1. In each case, name each reactant, predict each product, and identify each product.
2. What are the colors of each of the copper compounds?
3. Write the formula for each substance that is used and prepared in these reactions.
4. Write balanced equations for the reactions.

5. Chromium Reduction: Cold Orange to Hot Green

Several changes—from orange to green, from solution to solid gel, and from cool to hot—are all observed in a beaker when two solutions are mixed.

Procedure

1. Observe the temperature of the solutions of potassium dichromate and sodium bisulfite.
2. Place 50 mL of the dichromate solution in a 250-mL beaker and add an equal volume of bisulfite solution. Swirl the solutions to stir.
3. Make observations of the evidence that a chemical reaction has taken place.

Reactions

More than one reaction appears to take place in this system. The most significant reaction is the reduction of chromium and the subsequent formation of the hydrated hydroxide.



Perhaps the ions of sulfate and sodium are also incorporated in the gel to form an alum, $\text{KCr}(\text{SO}_4)_2$.

Solutions

1. The potassium dichromate, $\text{K}_2\text{Cr}_2\text{O}_7$, concentration is 0.5 M: Dissolve 147 g in 1 L of warm solution.
2. The sodium bisulfite, NaHSO_3 , concentration is 4.0 M: Dissolve 416 g in 1 L of solution.

Teaching Tips

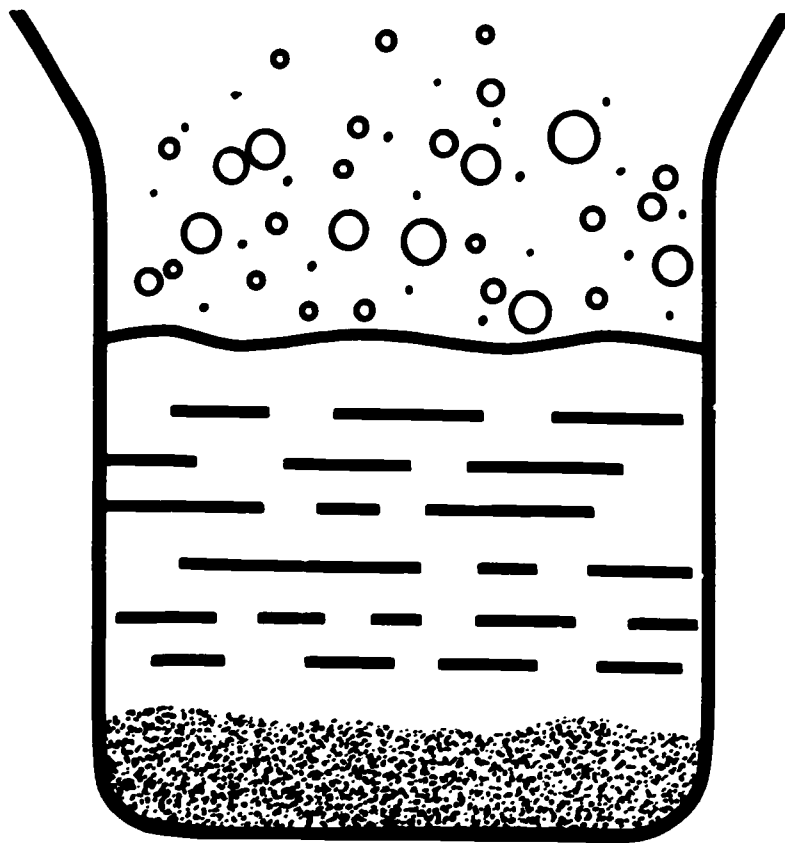
NOTES

1. This reaction is very dramatic as the clear orange solution becomes very dark green and turns into a solid light green gel while evolving enough heat to produce steam.
2. This demonstration acts as a good introduction to chemical reactions and equally well as an introduction to oxidation-reduction.
3. Notice that the concentration of bisulfite is 8 times that of the dichromate solution.
4. You may also use ammonium dichromate.
5. This system is not well described in the literature.
6. The chromium in a +3 oxidation state is nontoxic as compared with the suspected toxicity of +6 chromium. The gel can be easily disposed of in the wastebasket.
7. You might keep some of the gel exposed to the air for a few days to observe changes.

QUESTIONS FOR STUDENTS

1. Describe each observation as evidence that a chemical reaction was taking place.
2. What observation signals the change in oxidation state of the chromium? What is the change?
3. Looking at the concentrations of the solutions and the stoichiometry of the reaction, tell which reagent was in excess. (Sodium bisulfite.)
4. Describe the change in temperature.
5. Did you observe any evidence of gas formation? [SO_2 can sometimes be detected. $\text{H}_2\text{O}(\text{g})$ can be observed.]

Physical Changes



6. Nonadditivity of Volumes

Twenty-five milliliters of each of two miscible liquids is mixed in a narrow tube. The total volume is less than 50 mL.

Procedure

1. Place 25.00 mL of water in a 50-mL gas-collecting tube.
2. Add an additional 25.00 mL of acetone or ethyl alcohol by pouring the liquid carefully down the side of the tube until the 50.00-mL mark is reached.
3. Place your thumb over the mouth of the tube and invert the tube to mix the liquids.
4. Taking care not to lose any liquid on the thumb, examine the total volume.

Reactions

When these similar liquids are mixed, intermolecular forces act to cause the total volume to be less than that expected. A mixture of water and acetone is 4.5% less than the sum of the parts, and that of water and ethyl alcohol is 4% less. The expected reduction with the recommended volumes in this procedure, then, is about 2 mL.

Teaching Tips

NOTES

1. The greatest volume change with acetone in water is found when the acetone mole fraction is 0.25.
2. The explanation of this result requires the student to think about the individual particles of water and acetone.
3. Other measuring apparatus could be used, including large graduated cylinders and burets.
4. The smaller reduction in volume if one uses a test tube can be noticed by the pressure reduction on the thumb (suction). Half fill a small test tube with water. Fill the test tube to the top with ethyl alcohol. With the thumb covering the opening of the test tube, invert the test tube to mix the two liquids. The test tube will "stick" to the thumb.

QUESTIONS FOR STUDENTS

1. Try adding different volume fractions of the two liquids. What volume reduction is seen each time?
2. What are the formulas and the structures of the molecules of water and acetone or ethyl alcohol?
3. Suggest ways the two kinds of molecules might arrange themselves to cause the volume to be reduced.

7. The Mysterious Sunken Ice Cube

Two beakers of clear colorless liquid are displayed. In one beaker, an ice cube is floating. In the other beaker, an ice cube rests on the bottom.

Procedure

1. Before the students arrive, fill two medium-sized beakers three-fourths full—one with water and one with ethanol.
2. Place an ice cube in each beaker when you are ready to perform the demonstration.
3. Observe the positions of ice cubes.
4. After students make an initial observation, add 1 drop of food coloring to each beaker to improve visibility.

Solutions

Any liquid with a density of less than 0.92 g/cm^3 can be used.

Teaching Tips

NOTES

1. The density of water is 1.00 g/cm^3 . The density of ice is 0.92 g/cm^3 . The density of ethanol is 0.79 g/cm^3 .
2. This demonstration is a very simple way to interest students in the concept of density.
3. Display other examples of density differences as seen through buoyancy.
 - a. Use a battery jar of water in which you place a can of regular soda and a can of diet soda. The sugared version is more dense than water and will sink. The diet drink will float.
 - b. Half fill a large graduated cylinder with a concentrated salt solution. Carefully fill the cylinder with water. Slide in an egg. The egg will remain suspended at the interface.
 - c. Display equal-volume balloons filled with hydrogen, helium, laboratory gas, oxygen, nitrogen, air, and carbon dioxide.
4. Have students make some mass and volume measurements of the liquids to calculate density.
5. Ice floats on water because the molecules in the ice structure have crystallized through hydrogen bonding and form an "open" structure. No air molecules exist in the crystal structure as students often think. Air is trapped between ice crystals just as it is trapped in liquid water.
6. Icebergs are floating in seawater—density of 1.03 g/cm^3 . An iceberg is 11% exposed. Eighty-nine percent is under water.

QUESTIONS FOR STUDENTS

1. Explain why an ice cube floats in water.
2. Why do you think oil floats on water?
3. Give a stepwise procedure by which you could find the density of water, ethanol, or any other liquid.

8. Colorful Mixture Separations

Components of a mixture in a small beaker are separated by heating to sublime one (purple iodine), by dissolving one (yellow potassium chromate), and by leaving the last (dense sand) in the beaker.

Procedure

1. In a 100-mL beaker, mix together well with a glass rod a spatula tip of iodine (about $\frac{1}{16}$ tsp) and $\frac{1}{2}$ -1 tsp of clean white sand.
2. Add and mix in about $\frac{1}{4}$ tsp of potassium chromate crystals.
3. Observe the appearance of the mixture.
4. Put an ice cube in a small plastic bag (Baggie) in a small evaporating dish and place the dish on the beaker.
5. Place the beaker on a warm hot plate or over a VERY slow burning flame.
6. Observe the purple vapors of iodine gas and the recrystallized iodine on the underside of the cold dish.
7. When the iodine vapor is no longer visible in the beaker, remove the beaker from the hot plate and add 20-30 mL of water to the residue.
8. Stir and decant the potassium chromate solution; wash the residue several times with water.
9. Observe the three pure substances.

Reactions

These three substances have different physical properties, which allow easy separation.

1. The iodine sublimates—escapes from the solid directly to the gaseous phase.
2. The potassium chromate is soluble in water. Its solubility is 63 g/100 mL of water.
3. The silicon dioxide, sand, is insoluble.

Materials

1. Solid iodine, I_2 .
2. Solid potassium chromate, K_2CrO_4 .
3. Sand (silicon dioxide), SiO_2 .

Teaching Tips

NOTES

1. A small beaker works best. Iodine will crystallize on the inner walls of a larger one.
2. The iodine crystals that deposit on the underside of the evaporating dish are quite fascinating. Place the dish and a hand magnifier in a shallow box and pass them around. The crystals vary in size and are diamond-shaped.
3. Other colored solids may be tried; however, those solids with water of hydration will evolve water on heating.
4. If the mixture is thoroughly mixed, the students may not be able to predict how many components are present.

5. Iodine is obtained from deposits of sodium iodate, NaIO_3 , and from underground brines of mixed salts.
6. You might evaporate some of the potassium chromate solution.

QUESTIONS FOR STUDENTS

1. Name some common mixtures.
2. What property of iodine makes its separation possible?
3. Does iodine have a high or low vapor pressure?
4. Describe the crystals of pure iodine.
5. How could you obtain the potassium chromate as a pure solid?
6. What property of silicon dioxide allows it to be separated?

9. Surface Tension of Water: The Floating Needle

A needle floats on the surface of water in a beaker. One drop of soap solution is added, and the needle sinks to the bottom.

Procedure

1. Fill a large beaker almost full of water. Be sure that the beaker is clean.
2. Place a sewing needle on top of a small piece of tissue paper.
3. Place a large spatula or scoopula beneath the tissue paper and carefully lower it and the needle onto the water. The paper and needle will float. Soon, the paper will become soaked and sink, and the needle will be left floating on the surface.
4. After discussing the role of surface tension, add 1 drop of soap solution to the surface of the water.
5. Observe that the needle sinks in a few seconds.

Reaction

Surface tension of a liquid is the energy required to increase the surface area. This tension causes the surface to act as if it had a membrane, or "skin". This property is partly due to the fact that surface molecules are only pulled inward, whereas other molecules in the liquid are pulled equally in all directions.

The needle is denser than water, so if pushed under, the needle will sink. On the surface, however, energy would be required to increase the water surface enough for the needle to submerge, so it floats.

When detergent is added, the surface tension is drastically reduced and the needle sinks.

Teaching Tips

NOTES

1. Surface tension also allows some insects (water bugs, etc.) to walk on water, even though they are heavier than water; causes water to rise in a capillary tube; causes a drop of rain to be nearly spherical; and is one of the forces that causes water to rise in a plant.
2. The surface tension of water is $7.3 \times 10^{-2} \text{ J/m}^2$.
3. You may have difficulty placing the needle directly on the surface of the water.

QUESTIONS FOR STUDENTS

1. Why does the heavy needle float?
2. What is *surface tension*?
3. What does surface tension have to do with volume and area?
4. What other effects can you think of that are due to surface tension?

10. Liquid Nitrogen

Liquid nitrogen is used to "quick freeze" several items and perform other demonstrations that require a very low temperature.

Procedure

CAUTION: YOU MUST BE EXTREMELY CAREFUL WHEN HANDLING LIQUID NITROGEN. THE LIQUID SHOULD ONLY BE PLACED IN A DEWAR FLASK OR SIMILARLY INSULATED CONTAINER. HANDLE THE CONTAINER ONLY WITH INSULATED GLOVES AND DO NOT ALLOW IT TO COME IN CONTACT WITH THE SKIN.

QUICK FREEZING

1. Fill a widemouthed Dewar flask with liquid nitrogen.
2. Holding a fresh carnation by its stem, insert the flower in the flask. After about 10 s, remove the flower and crumple it in your hand.
3. Holding a banana by one end, insert it in the flask. After about 2 min, remove the banana and show that it is frozen by using it as a hammer to drive a small nail (with a large head) into a soft board.
4. Place about a dozen grapes in a skillet or similar container. Carefully pour liquid nitrogen into the pan until it is about half full. After a minute or so, rattle the grapes to show that they are frozen. Pour the extra liquid in the pan on the floor, away from any of the students.
5. Using crucible tongs, hold a small HOLLOW rubber ball in the flask of liquid nitrogen. After about 2 min, remove the ball and throw it against a wall (away from the students), where it will shatter.

TEMPERATURE AND VOLUME EFFECTS

1. Blow up a long balloon that is wide enough to fit into the mouth of the Dewar flask.
2. Carefully insert the balloon into the flask. Continue to insert the balloon as it "shrinks" in the cold liquid.
3. When the entire balloon has been cooled, place it on the table top to observe as it warms and expands to its original volume. (You can also pour liquid nitrogen on top of a round, inflated balloon for the same effect.)
4. Carefully pour about 50 mL of liquid nitrogen into a 250-mL Erlenmeyer flask. With someone securely holding the flask, place a deflated balloon over the mouth of the flask. Observe the balloon for several minutes. Notice also that water vapor from the air condenses and freezes on the flask.

Teaching Tips

NOTES

1. The temperature of liquid nitrogen is -196°C .
2. Liquid nitrogen is made by condensing air. Its boiling point is less than that of oxygen, so liquid nitrogen evaporates first from liquid air (fractional distillation).
3. Liquid nitrogen is used as a freezing agent in the food-processing industry.
4. The only danger in using liquid nitrogen is its very low temperature. Unlike liquid oxygen, liquid nitrogen is unreactive.

5. To obtain liquid nitrogen, contact a college or university in your area. Also, contact local dermatologists—they often use liquid nitrogen to remove certain growths from the skin. Check the Yellow Pages.

QUESTIONS FOR STUDENTS

1. What do you think is the chief substance that is frozen when the flower, the grapes, and the banana freeze? (Water, as in any cell.)
2. If the ball is already a solid, how can you account for the change in its properties when it becomes as cold as the liquid nitrogen?
3. What seems to be the comparison of the volume of liquid nitrogen to the volume of that same nitrogen sample when it is a gas in the balloon?
4. Explain how nitrogen can be separated from air (80% nitrogen).

11. The Cartesian Diver: An Application of Boyle's Law

A large plastic soft drink bottle is passed around the class. The bottle is filled with water and has a small dropper at the top. As the sides of the bottle are squeezed, the dropper sinks to the bottom of the bottle. When the pressure is released, the dropper rises to the surface again.

Procedure

1. Remove the label from the outside of a large plastic soft drink bottle.
2. Fill the bottle to the top with water.
3. Place an empty medicine dropper in the bottle. Be sure that the bottle is full to the top.
4. Put the cap on the bottle and tighten.
5. Gently press the sides of the plastic container. Notice that the water level will rise in the dropper and the dropper will slowly begin to sink.
6. Release the pressure, and the water level in the dropper will become lower and the dropper will rise again.
7. Pass the "Cartesian diver" around your class.

Reaction

According to Boyle's law, increasing pressure on a gas will decrease its volume. Pressing on the sides of the plastic bottle increases the pressure on the water and forces it into the dropper. This increase in water volume reduces the volume of air in the dropper. The dropper is now heavier and it sinks. When the pressure is released, the opposite effect occurs.

Teaching Tips

NOTES

1. Increasing or decreasing the amount of water in the dropper decreases or increases its *buoyancy*; thus, the dropper sinks or rises.
2. Try using other containers, including glass containers.
3. Students can easily SEE the change in gas volume in the dropper in this demonstration.
4. It is important that students realize that the AMOUNT of air in the dropper is not changing, only its VOLUME.
5. "Cartesian" pertains to the work of Descartes, the French philosopher (1596-1650).

QUESTIONS FOR STUDENTS

1. Describe the steps needed to cause the dropper to sink.
2. Discuss what happens to the air in the dropper.
3. Would this demonstration work with a liquid other than water? Try it!

12. Charles' Law: The Relationship between Volume and Temperature of a Gas

Balloons are inflated to the same size and tied. One balloon is placed on the surface of a large beaker of ice water. Another balloon is placed on the surface of a beaker of hot water. At the decreased temperature, the balloon shrinks. At the increased temperature, the other balloon becomes larger.

Procedure

1. Blow up three balloons—not fully—to the same size and tie securely.
2. Leave one balloon for reference. Place one balloon on the surface of a large beaker (or pan) of ice water.
3. Place the third balloon on the surface of a large beaker of hot water.
4. Observe the effect of temperature on the volume of the gas in the balloons.
5. Remove both balloons and allow them to come to room temperature.

Reactions

As the temperature decreases, the average kinetic energy of the gas inside the balloon decreases, which decreases the pressure inside the balloon. When the balloon is heated, the volume increases as the internal pressure increases.

Teaching Tips

NOTES

1. Try various hot and cold temperatures to get maximum effects.
2. Jacques Alexandre Charles made the first observation of gases at different temperatures in 1787.
3. Putting his experience to practical use, Charles was one of the first hot-air balloonists. In 1783, he made the second ascent in a hydrogen-filled balloon carrying passengers. When he landed, his balloon was destroyed by terrified, pitchfork-wielding peasants.
4. The change in gas volume is essentially the same for each Celsius degree of temperature change, approximately $\frac{1}{273}$ of the volume that gas would have at 0 °C.

QUESTIONS FOR STUDENTS

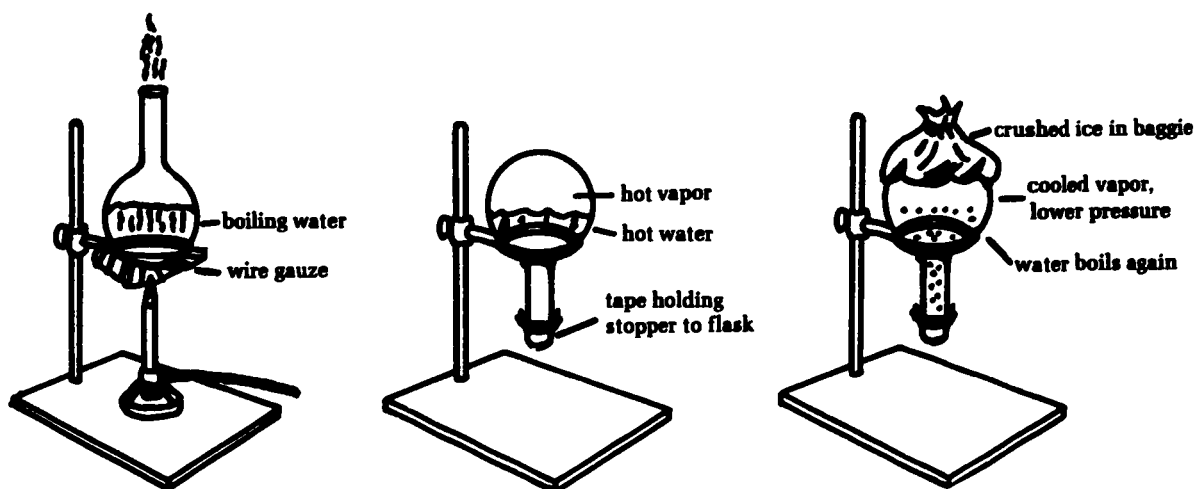
1. What is Charles' law?
2. Why do the balloons behave as they do in hot and cold water?
3. Why do the balloons return to "normal" at room temperature?

13. Boiling at Reduced Pressure

A flask of boiling water is stoppered and inverted. The water and vapor in the flask are cooled by crushed ice, and the water boils at a low temperature.

Procedure

1. Place the flask on wire gauze as shown in the diagram.
2. Fill the flask approximately half full of water.
3. Heat the flask until the water boils and steam is produced.
4. Remove the burner.
5. Firmly insert the stopper in the flask. Wrap masking tape around the stopper to be sure that it is held in place.
6. Remove the wire gauze and place the flask neck down through the ring. Notice that the water is hot, but NOT boiling.
7. Place a closed baggie full of crushed ice directly on top of the inverted flask.
8. Observe.



Reactions

A liquid *boils* when the molecules gain enough energy to escape the surface of the liquid and produce a pressure equal to the pressure of the atmosphere or vapor above the liquid. Water must be heated to 100 °C at normal atmospheric pressure before the molecules have enough energy. When the pressure above the water is much lower than atmospheric pressure, as when the condensed steam leaves a low pressure above the water in the flask, the water will then boil when very little external energy is added to the system. The water boils when it is quite cool.

Materials

1. The flask must be ROUND-bottomed.
2. Use a baggie for the crushed ice.

Teaching Tips

NOTES

1. Be sure to tape the stopper into the neck of the flask. If the stopper should come out, hot water will spill over the desk.
2. Encourage students to measure the temperature of the water as it "boils".
3. When a substance has a high vapor pressure, its molecules will break their attraction for other molecules in the liquid and enter the atmosphere as a gas. We can detect substances with high vapor pressures by smelling them. Examples are perfumes, ether, and so forth.

QUESTIONS FOR STUDENTS

1. Do you think that liquids would boil at a higher or lower temperature at, say, 20,000 ft above sea level?
2. What effect did adding the ice pack have on the boiling point?
3. Can you make warm water boil by this method? Try it! What is the lowest temperature at which water will boil when the ice pack is added? Try it!

14. **The Collapsing Plastic Soft Drink Bottle and Soft Drink Can**

The Collapsing Soft Drink Bottle

A 2- or 3-L empty plastic soft drink container is filled to the top with hot water. The water is poured out of the bottle, and the empty bottle is immediately capped. In a few seconds, the bottle begins to collapse. The usual effect is the production of a four-sided container!

Procedure

1. Obtain an empty 2- or 3-L plastic soft drink bottle.
2. Fill the bottle to the top with hot (almost boiling) water. **CAREFUL!**
3. Immediately pour the water from the bottle and quickly screw the cap tightly on the bottle.
4. Place the bottle in full view of your class.
5. Observe.

Reaction

The hot water inside the bottle heats the air in the bottle, which greatly reduces the pressure inside the container, when the bottle is capped and the air cooled. The pressure outside is greater than the pressure inside, which causes the bottle to collapse.

The Collapsing Soft Drink Can

An aluminum soft drink can is heated until it is filled with water vapor. The can is suddenly removed from the heat and inverted in a thin layer of water in a trough. The can collapses suddenly and dramatically.

Procedure

1. Place about 5 mL of water in an empty aluminum soft drink can.
2. Boil the water in the can vigorously for several minutes, until the can is filled with water vapor.
3. Carefully (use tongs) and quickly invert the can and place it, top down, into a trough containing about 1 in. of water.
4. Observe!

Reaction

The water cools the can and condenses the steam inside. The sudden reduction in pressure in the can causes the can to collapse.

Teaching Tips

NOTES

1. These demonstrations are nice because they are simple and quickly performed.
2. The usual effect is for the round bottle to become four-sided!
3. If your bottle does not collapse, try hotter water. If your bottle softens, try cooler water.
4. Be sure to use seamless, aluminum cans.
5. The hotter the can, the more dramatic the results.
6. Use a burner or arrange several cans on a hot plate so the effect may be observed several times.

QUESTIONS FOR STUDENTS

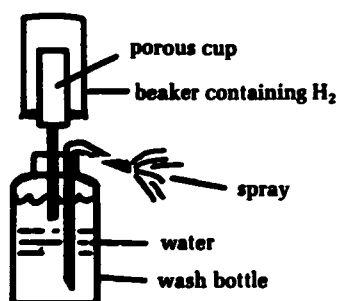
1. Compare the gas pressure inside with the pressure outside the can and bottle at the point of collapsing.
2. Compare the numbers of gas molecules inside the can before and after the steam condenses.
3. What causes the warm plastic bottle to collapse?

15. The Automatic Water Fountain: Hydrogen Effusion

A plastic wash bottle is prepared with a porous clay cylinder connected to the top (see the diagram). When a beaker of hydrogen gas is inverted over the cylinder, water from the wash bottle is forced out as a spray.

Procedure

1. Using a short piece of glass tubing, insert a stoppered porous clay cup into a one-hole stopper in the mouth of a plastic wash bottle. (The wash bottle must have a tube extending to the bottom.)
2. Fill the bottle half full of water and add 1 or 2 drops of food coloring for effect.
3. Fill a large beaker with hydrogen (see the diagram).
4. Invert the beaker over the porous cup and hold the beaker in place.
5. Observe the spray as the water is forced out of the wash bottle.



Reactions

1. The porous cup allows small molecules to enter and leave (effuse). Because the rate of effusion of gases is inversely proportional to their molar masses (Graham's law), hydrogen effuses more rapidly than other gases, including oxygen and nitrogen, which are inside the clay cup.
2. Because hydrogen enters the clay cup, pressure in the cup increases, and this increase results in increased pressure in the wash bottle. This increased pressure forces the water out of the bottle as a spray.

Teaching Tips

NOTES

1. You can, of course, construct a wash bottle from a flask to serve the same purpose.
2. For best results, obtain hydrogen gas from a tank. You may also prepare hydrogen by reacting zinc and hydrochloric acid. Collect the hydrogen by displacement of water.
3. *Effusion*, a gas escaping from a container through small openings (like the porous cup), is often confused with *diffusion*, the mixing of one gas with another. Graham's law also applies to diffusion.

QUESTIONS FOR STUDENTS

1. Which would effuse the fastest: hydrogen, helium, or nitrogen?
2. At the same temperature, which molecule has the highest average velocity: hydrogen, nitrogen, or water vapor?
3. What is the difference between *effusion* and *diffusion*?

16. Eutectic Solidification

Small cubes of two metals are placed in glycerol heated to about 180 °C. This temperature is below the melting point of either metal. When the metals are pressed together, a liquid forms between them and they become fused.

Procedure

1. Heat about 100 mL of glycerol in a 250-mL beaker to about 180 °C. Maintain a constant temperature.
2. Clean the surface of a cube of bismuth and a cube of tin with alcohol. If the faces are not smooth, sand them with sandpaper or flatten them with a hammer.
3. Place the two metal pieces in the beaker of glycerol. The metal pieces should not be in contact. Heat the beaker for 1-2 min to allow the metals to reach bath temperature.
4. Poke at each piece with a stirring rod to convince students that the metals are still solid.
5. Using two stirring rods, push the two metals together and hold them together until signs of liquid can be seen between them. (This step will probably take 30 s-1 min).
6. Allow the glycerol to cool and remove the fused metal samples.

Reaction

1. The melting point of Sn is 232 °C; that of Bi is 271 °C.
2. When the two metals are in contact, diffusion occurs between them. The mixture of Sn and Bi (alloy) at the surface has a lower melting point (139 °C) than the melting point of either metal alone. This value is the EUTECTIC point for 57% Bi.

Teaching Tips

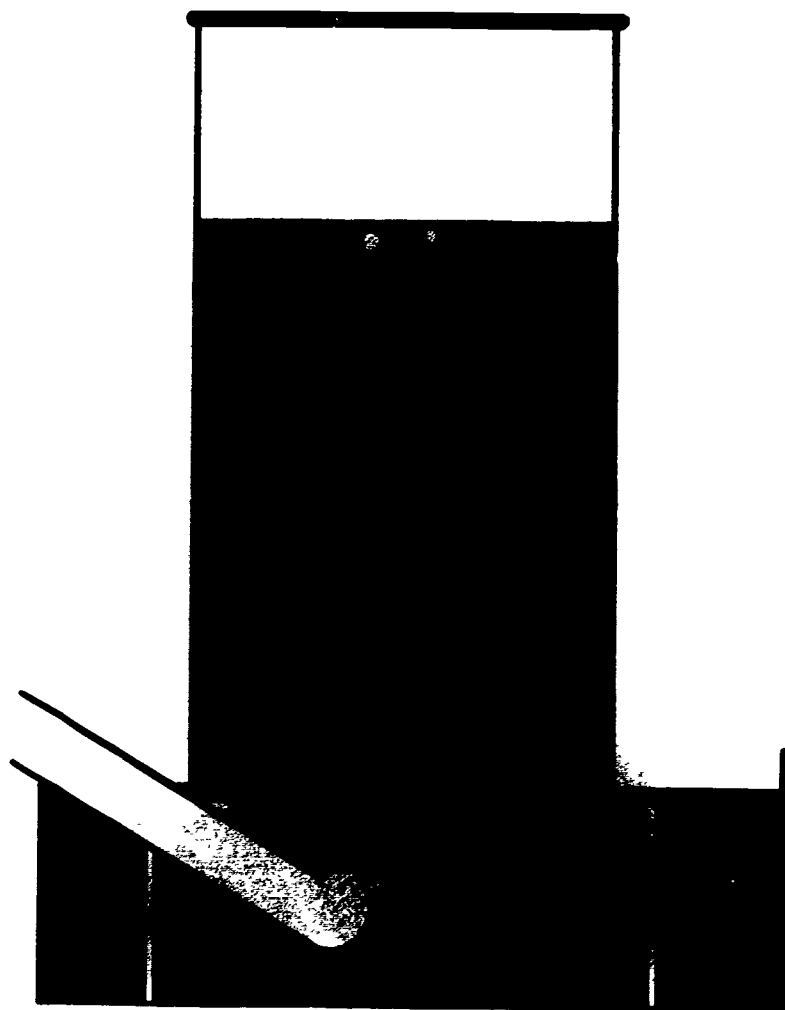
NOTES

1. Alloys such as this are used for solder, to fuse metals, and in semiconductors for gold bonding of silicon to metal headers.
2. This process can be compared to lowering the freezing point of water with rock salt.
3. Have your students look up the eutectic point of other common alloys (bronze, brass, etc.).

QUESTIONS FOR STUDENTS

1. Why is this reaction carried out in glycerol rather than in water?
2. What is a *eutectic* point?
3. How does the eutectic point vary from the melting point of its two solids?
4. What practical use do eutectic mixtures have?

Reactions Involving Gases



17. Making Hydrogen Gas from an Acid and a Base

Two flasks are fitted with balloons. One flask contains hydrochloric acid and aluminum metal. The other flask contains sodium hydroxide and aluminum. As the reactions proceed, balloons on both flasks become inflated because of the production of hydrogen gas.

Procedure

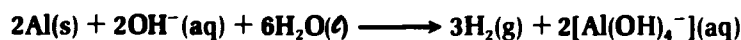
1. Place 20–30 mL of 3 M hydrochloric acid in one 250-mL flask.
2. Place the same volume of 3 M sodium hydroxide in another 250-mL flask.
3. Have two large balloons and a student assistant ready.
4. Drop a loose wad of aluminum foil into each flask and stretch a balloon over each. Use the same size piece of foil in each flask.
5. Watch for evidence of a reaction, relative rates of reactions, differences in the products, and the presence of any excess reactant.
6. When the balloons are inflated, carefully remove them, tie the ends, and attach the balloons to the end of the demonstration table.
7. Using a meter stick with a candle taped to one end, carefully light each balloon. Compare the effects.

Reactions

1. Active metals, such as aluminum, react with acids to release hydrogen gas.



2. Aluminum also dissolves in a strong base to form hydrogen gas and tetrahydroxaluminate(III) ion.



3. The strong acid and strong base dissolve the oxide film on aluminum; the aluminum thus comes into direct contact with the acid or base to produce hydrogen.

Solutions

1. The hydrochloric acid concentration is 3 M: See Appendix 2.
2. The sodium hydroxide concentration is 3 M: Dissolve 120 g of NaOH in 1 L of solution.

Teaching Tips

NOTES

1. Aluminum, zinc, and gallium are among the few metals that will release hydrogen from a base. The reaction with gallium is similar to that of aluminum; this reaction produces sodium gallate(III). Zinc produces sodium zincate. You might repeat the demonstration using zinc.
2. You cannot use nitric acid because it forms a coating of aluminum oxide on the metal.

3. Buy the larger party balloons. The balloons in regular packages are too small.
4. You might like to collect the hydrogen produced by these reactions in the conventional manner (by water displacement) and test for the properties of hydrogen gas.

QUESTIONS FOR STUDENTS

1. Show how hydrogen gas can be produced from both a base and an acid.
2. Compare the rates of reaction in the two flasks.
3. Why is aluminum left unreacted in the flask with hydrochloric acid? (Three times the amount of acid is required than that of sodium hydroxide per mole of aluminum.)
4. What difference do you notice in the flasks as the reactions proceed?
5. What evidence existed that hydrogen gas was produced?
6. What other metals will undergo these reactions?

18. Dancing Mothballs and Dancing Spaghetti

Mothballs placed in a large graduated cylinder continually rise to the top and then slowly sink to the bottom. Pieces of uncooked spaghetti do the same in another cylinder.

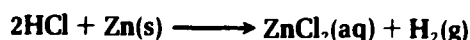
Dancing Mothballs

Procedure

1. Fill a large graduated cylinder with water. For best results, use a 1-L cylinder.
2. Add 10 mL of concentrated HCl. ADD THE ACID SLOWLY, ALLOWING IT RUN DOWN THE SIDES OF THE CYLINDER.
3. Carefully drop in six to eight small pieces of new mossy zinc.
4. Drop in three to four mothballs. (Only *para*-dichlorobenzene mothballs work!)
5. Observe. After a few minutes, the balls will mysteriously rise to the top of the cylinder and then sink to the bottom.

Reactions

1. The HCl reacts with Zn to produce hydrogen gas.



2. The bubbles of hydrogen attach to the mothballs, which causes them to rise to the surface.
3. When they reach the surface, the hydrogen escapes, and the mothballs fall to the bottom

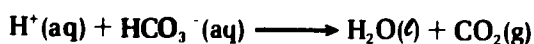
Dancing Spaghetti

Procedure

1. Almost fill a large graduate cylinder or quart jar with water.
2. Place 2 Tbs of baking soda (NaHCO_3) in the solution.
3. Add a handful of broken spaghetti.
4. Slowly add up to 100 mL of vinegar until the spaghetti begins to "dance."

Reaction

Acetic acid in vinegar reacts with baking soda to produce carbon dioxide gas, which causes an effect like that of hydrogen.



Teaching Tips

NOTES

1. Practice to get just the right amount of zinc or baking soda and acid.
2. A 100-mL cylinder is too small. The mothballs and spaghetti cannot "pass each other" on the way up and down.

3. Do not use powdered zinc.
4. Come up with your own variation of this demonstration. Try spaghetti and a glass of soda pop!

QUESTIONS FOR STUDENTS

1. Describe the difference between a rising mothball (or piece of spaghetti) and a falling mothball (or piece of spaghetti).
2. Does it matter what gas is used when this effect is produced?
3. What variables can be changed to cause the mothballs to rise at a different rate?

19. Producing Methane Gas

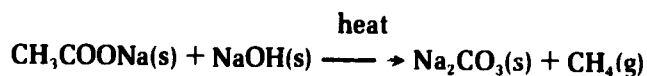
Two solid chemicals are heated in a test tube with a delivery tube leading to a water trough. When the tube is heated, methane gas is produced and is collected in test tubes by displacement of water.

Procedure

1. Mix approximately 10 g of sodium acetate with the same amount of sodium hydroxide.
2. Immediately place the mixture in a large, DRY, test tube fitted with a stopper and a gas delivery tube.
3. Prepare to collect a gas from the reaction by inverting several test tubes filled with water in a large water trough.
4. Carefully heat the mixture in a test tube. Note the formation of a gas.
5. Collect several test tubes of gas. **SHUT OFF THE BUNSEN BURNER.**
6. Test the gas by carefully inserting a lighted wooden splint into the test tube. **CAUTION: POINT THE TUBE AWAY FROM STUDENTS WHEN LIGHTING IT.**

Reactions

1. Production of methane gas:



2. Combustion of methane gas:



Materials

You must use dry, anhydrous sodium acetate and dry sodium hydroxide.

Teaching Tips

NOTES

1. Various reactions for the combustion of methane are possible, depending upon the number of moles of oxygen available. Try other reactions.
2. Methane is lighter than air.
3. Both solids used in this demonstration are very hygroscopic, so they must be placed in the test tube and heated quickly.
4. The test tube must be clean and dry.

QUESTIONS FOR STUDENTS

1. Write the reaction for the production of methane in this demonstration.
2. Considering the density of methane gas, should the test tubes of methane be held upright or inverted when the lighted splint is used?
3. How could you show that methane gas is lighter than air?

20. A Hand-Held Reaction: Production of Ammonia Gas

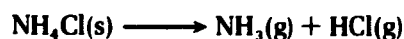
Small amounts of two white solids are placed in the palm of the hand. When the solids are rubbed together, ammonia gas is produced and can be easily detected.

Procedure

1. If your hand is clean and dry, place an amount of ammonium chloride that would cover a dime in the palm.
2. Add an equal amount of calcium hydroxide on top of the ammonium chloride.
3. Rub the two solids together briskly and check for the odor of ammonia gas.

Reactions

1. The small amount of friction between the two solids is sufficient to begin the decomposition of ammonium chloride.



2. The calcium hydroxide reacts with the hydrogen chloride gas and prevents it from being detected.



Teaching Tips

NOTES

You should, of course, wash your hands thoroughly after this demonstration!

2. This demonstration is a good example of a reaction between two solids. Another good example is to mix solid potassium iodide and lead nitrate to form yellow lead iodide.
3. If you would like to confirm the presence of HCl (acid) and ammonia (base), see demonstration 21.

QUESTIONS FOR STUDENTS

1. Write the equation for the reaction to produce ammonia.
2. What does the calcium hydroxide do in this reaction?
3. Classify these reactions as exothermic or endothermic.

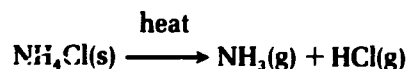
21. Producing Two Gases from Ammonium Chloride

Solid ammonium chloride, when heated, decomposes to produce both ammonia gas and hydrogen chloride gas. The basic and acidic properties of these gases are detected with litmus paper.

Procedure

1. Add enough solid ammonium chloride to fill a large, clean, dry test tube about one-fourth full.
2. Insert a plug of glass wool in the mouth of the test tube.
3. Have moist strips of red and blue litmus paper available to test for the presence of ammonia and hydrogen chloride gases.
4. Holding the test tube at a slight angle, heat the ammonium chloride.
5. While heating the ammonium chloride, hold moist strips of red and blue litmus paper over the mouth of the test tube.
6. Continue heating until the red litmus paper turns blue, which indicates the presence of a base (ammonia).
7. With additional heating, the blue litmus paper will turn red, because of the production of hydrogen chloride (acid).
8. Notice, also, the condensation of ammonium chloride on the cooler part of the test tube.

Reaction



The ammonia is detected first because, even though both gases are produced simultaneously, it is lighter and leaves the test tube first. HCl is about twice as heavy as ammonia, travels more slowly, and is detected later.

Teaching Tips

NOTES

1. The glass wool acts as a barrier that separates the lighter, faster moving ammonia gas and traps vaporized ammonium chloride; the ammonium chloride therefore condenses in the test tube.
2. Vary the amounts of ammonium chloride for maximum effect.
3. Notice that this reaction is reversible: cooling causes solid ammonium chloride to re-form.

QUESTIONS FOR STUDENTS

1. How did ammonia gas and hydrogen chloride gas affect the litmus paper?
2. Why was it necessary to use "moist" litmus paper?
3. What properties of the gas molecules enable them to be detected in this demonstration?

22. Fluidity of Gases

Two gases are produced and "poured" down a trough onto a candle flame. One gas causes the flame of the candle to go out. The other gas ignites and produces a flame that travels back up the trough into the container where it was produced.

Procedure

Producing Carbon Dioxide Gas

1. Prepare a U-shaped trough about 12 in. long by shaping heavy-duty aluminum foil or trimming an aluminum pastry pan.
2. Fasten the metal trough so that it will empty directly above a flame of a small candle when the trough is placed at a 45° angle.
3. Provide a stand at the top of the trough, where a 250-mL beaker can be placed on its side and securely fastened.
4. Place a small piece of dry ice in the beaker. Add about 5 mL of water. Light the candle at the bottom end of the metal trough. Place the beaker on its side, so that the carbon dioxide gas flows out and down the trough. Secure the beaker in position.
5. Observe.

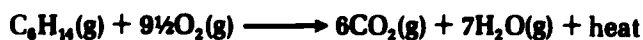
Naphtha, a Volatile Hydrocarbon

THIS DEMONSTRATION WILL PRODUCE A FLAME. YOU MUST USE A METAL TROUGH AND A PYREX BEAKER AND HAVE NO COMBUSTIBLE MATERIAL NEARBY.

1. Set up the aluminum foil assembly as described under Producing Carbon Dioxide Gas.
2. Do not light the candle yet.
3. Place a small squirt of cigarette lighter fluid in the beaker.
4. Immediately place the beaker on its side and secure it in place. YOU WILL NOT BE ABLE TO MOVE THE BEAKER, BECAUSE THE NAPHTHA WILL BE BURNING LATER.
5. Light the candle quickly.
6. Observe, but from a safe distance, as the candle ignites the naphtha gas and produces a flame that travels up the trough and into the beaker where the remaining naphtha burns.
7. Do not attempt to remove the beaker until the flame is out and the beaker has cooled.

Reactions

1. Because the density of carbon dioxide (1.98 g/L) is much greater than that of air (1.29 g/L), carbon dioxide stays in the trough where it flows out the bottom and extinguishes the candle.
2. Naphtha, a petroleum product that vaporizes readily at room temperature, is also more dense than air. However, this vapor is flammable, and it ignites when it flows down to the candle flame. Naphtha is a mixture of hexanes, which can be represented generally as



Materials

1. Dry ice: You can often obtain small amounts of dry ice from ice cream parlors.
2. Naphtha: Cigarette lighter fluid (Ronsonol, for example) is naphtha.

Teaching Tips

1. If you do not have dry ice, you can easily generate CO_2 gas by adding 5–10 mL of dilute HCl to about 1 spoonful of calcium carbonate in the beaker.



You can also use vinegar (5% acetic acid) and sodium bicarbonate, NaHCO_3 .

2. There is no danger of the naphtha exploding, but it will burn. You must secure the beaker to prevent it from rolling off the platform.
3. You can use a tin can, but students will not be able to see the flame inside the container.
4. Try both of these demonstrations with the lights dimmed.
5. The fluidity of ether vapor has been a frequent cause of fires in chemical laboratories.
6. Naphtha is a petroleum distillate that is collected between 70 and 80 °C. It is a mixture of hexanes and perhaps some heptane. It is also called *benzine*, not to be confused with benzene.

QUESTIONS FOR STUDENTS

1. What is meant by *fluidity* of a gas?
2. What can you say about the density of naphtha vapor? (It is greater than air.)
3. Which of these gases would you expect to show fluidity? The density is given for each: ammonia, 0.77 g/L; chlorine, 3.21 g/L; helium, 0.18 g/L; and hydrogen chloride, 1.64 g/L. (Only those gases with a density greater than that of air, 1.29 g/L).

23. The "Aladdin's Lamp" Reaction

Removal of the stopper from an "Aladdin's lamp" (any opaque bottle or container) results in the delayed evolution of a large cloud of water vapor somewhat like the mystical "jinni".

Procedure

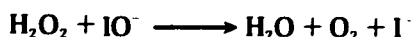
1. Obtain a large opaque bottle and a stopper or cap to fit.
2. Add 30-50 mL of 30% hydrogen peroxide to the bottle. CAREFUL: WEAR RUBBER GLOVES WHEN HANDLING 30% HYDROGEN PEROXIDE. CONTACT WITH THE SKIN MAY CAUSE BURNS.
3. Remove the tea from a tea bag and refill it with $\frac{1}{4}$ tsp of solid potassium iodide. Tie the tea bag securely with its string.
4. Place the tea bag inside the neck of the bottle. While holding the string, insert the stopper to hold the tea bag in place.
5. When ready to perform the demonstration, remove the stopper and allow the tea bag to fall into the hydrogen peroxide. Be sure to point the neck of the bottle up and in a safe direction.
6. Within a few seconds, an exothermic reaction in the bottle will produce a large amount of oxygen gas and water vapor that will form an impressive cloud as it escapes from the bottle.

Reactions

1. The iodide ion, I^- , from KI catalyzes the decomposition of hydrogen peroxide in an exothermic reaction to produce oxygen gas and water vapor.



2. The reaction is believed to occur in two steps with the IO^- ion as a reaction intermediate.



Solutions

1. Hydrogen peroxide is a 30% solution. The drugstore variety, a 3% solution, will not work to produce the effect of this reaction.
2. Potassium iodide: Use a fine, granular form of KI.

Teaching Tips

NOTES

1. Other solids, including manganese dioxide, may be used to catalyze this reaction.
2. Catalase, an enzyme in the blood, reacts with the hydrogen peroxide applied to a wound to produce oxygen and give a bubbly effect.
3. This reaction is a good example of catalysis, decomposition, exothermic reactions, and reaction mechanisms.

4. A 100% hydrogen peroxide solution is a stable, syrupy liquid with a pale blue color. However, it may decompose with explosive force in the presence of even small quantities of impurities.

QUESTIONS FOR STUDENTS

1. Which is more stable, the hydrogen peroxide molecule or the water molecule? Why?
2. Show that the catalyst is not "consumed" in this reaction.
3. What is the structure of hydrogen peroxide? ($\text{H}-\text{O}-\text{O}-\text{H}$)

24. Sparkler in Pure Oxygen

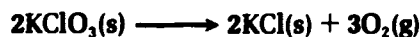
Oxygen is generated quickly and collected by water displacement. A "sparkler" is made by adding burning steel wool to the oxygen in a flask.

Procedure

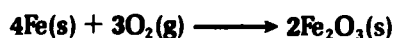
1. In a beaker, carefully mix a few teaspoonfuls of potassium chlorate and $\frac{1}{4}$ spoonful of manganese dioxide. DO NOT GRIND. Mix by shaking the beaker.
2. Place the mixture in a large clean Pyrex test tube fitted with a rubber stopper with a glass bend and length of tubing.
3. Assemble a container of water and an inverted 500-mL Pyrex flask filled with water for gas collection.
4. Heat the mixture in the test tube and observe the evolution of oxygen gas.
5. After the gas-air combination has evolved from the test tube for 20-30 s, begin collecting the oxygen in the flask by water displacement.
6. When the flask is full of gas, cover it with a glass plate and remove it from the water.
7. Remove the tubing from the water before you stop heating the test tube.
8. Holding a piece of steel wool in a burner flame, ignite the steel wool and immediately drop it into the flask.
9. Compare the rate of combustion of iron in air and in oxygen.

Reactions

1. Generation of oxygen is a typical decomposition reaction.



2. Burning of iron is a typical composition (synthesis) reaction and produces iron(III) oxide.



Materials

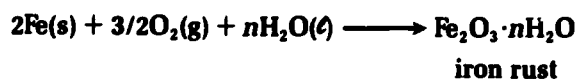
1. Potassium chlorate, KClO_3 .
2. Manganese dioxide, MnO_2 .
3. Steel wool.

Teaching Tips

NOTES

1. It is worthwhile for students to see the generation of oxygen although this procedure is considered too hazardous for students to perform. The chlorate is a strong oxidizing agent and may explode in the presence of any organic material. Do not mix the chlorate with anything other than the manganese dioxide.
2. When the oxygen is collected, remove the tubing from the water before the test tube begins to cool so cold water does not enter the hot tube and crack it.
3. Oxygen may also be generated by adding MnO_2 to 3% hydrogen peroxide and collecting the oxygen similarly.

4. Slow combination of iron with oxygen produces iron "rust".



QUESTIONS FOR STUDENTS

1. What class of chemical reaction is the one producing oxygen?
2. Write the equation.
3. What class of chemical reaction is the one using oxygen?
4. Account for the difference in rates of burning in oxygen and in air.

25. A Simple Reaction To Produce Foam

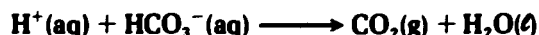
When two clear solutions are mixed, a large amount of foam is produced.

Procedure

1. Place 50 mL of solution A (vinegar and detergent) in a large beaker.
2. Place 50 mL of solution B (water and baking soda) in a 500-mL graduate or other tall container.
3. Pour solution A from the beaker into solution B in the tall cylinder.
4. Observe!

Reaction

This foam is produced by the production of carbon dioxide gas in a detergent solution when acetic acid (HCH_3COO) in vinegar reacts with baking soda (NaHCO_3).



Solutions

1. Solution A: Place about 1 Tbs of laundry detergent in 50 mL of white vinegar (acetic acid).
2. Solution B: Dissolve about 1 Tbs of baking soda (sodium bicarbonate) in 50 mL of water.

Teaching Tips

NOTES

1. Vinegar is a 5% acetic acid solution.
2. This foam is easy to produce and is suitable for elementary and junior high school students.
3. Ask students to suggest other ways CO_2 could be trapped to make foam. (Egg white, etc.)
4. Liquid detergent works well because it produces a clear solution. However, granular Tide produces copious foam. Consider having students experiment with the foaming character of various detergents by using this reaction.

QUESTIONS FOR STUDENTS

1. What is the foam made of? (CO_2 trapped in detergent bubbles.)
2. Design an experiment to test the foaming ability of various detergents and soaps. Be sure to control all variables except the brand of detergent or soap.

26. A Gas Evolution Oscillator

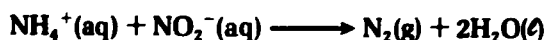
When two solutions are mixed in a large test tube, intermittent frothing occurs as the oscillating reaction produces nitrogen gas during one step of the reaction.

Procedure

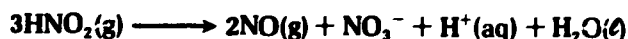
1. Add 10 mL of solution A to a large test tube (25 × 200 mm).
2. Add 10 mL of solution B.
3. Hold the test tube at an angle and observe for intermittent gas evolution.

Reactions

1. The reaction is the decomposition of aqueous ammonium nitrite.



2. A side reaction produces NO(g), which reacts with the oxygen in the air to produce NO₂(g).



Solutions

1. Solution A: Dissolve 26.4 g of ammonium sulfate in 100 mL of 0.2 M H₂SO₄.
2. Solution B: Dissolve 27.6 g of sodium nitrite, NaNO₂, in 100 mL of water. Add 1.0 g of potassium perchlorate, KClO₄.

Teaching Tips

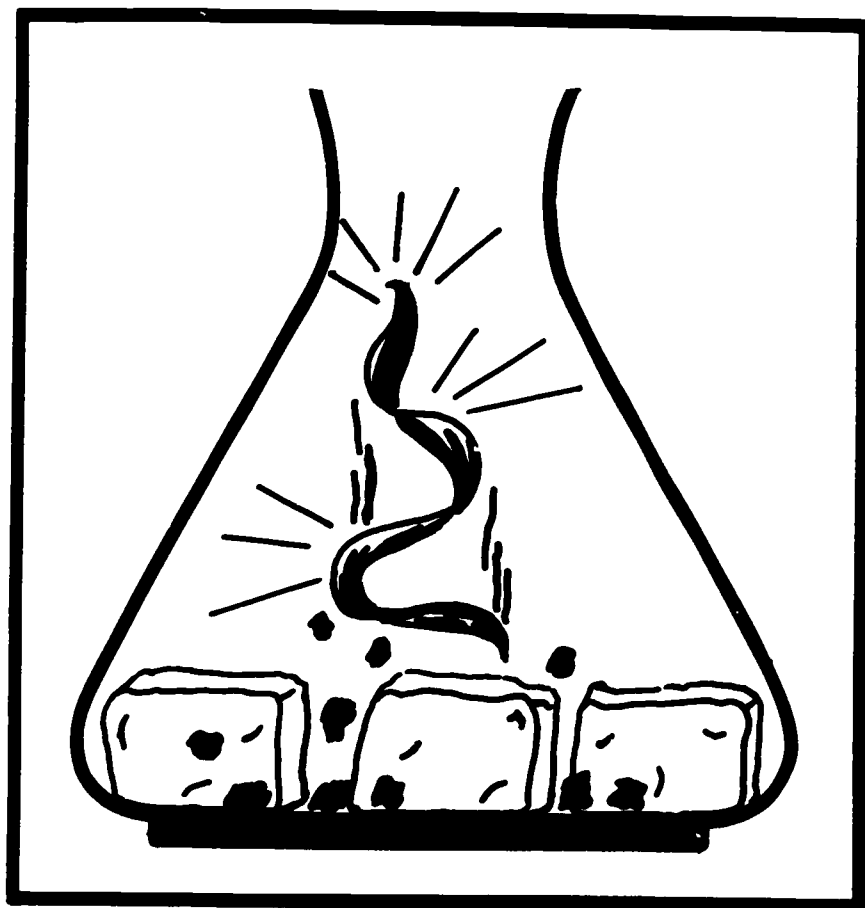
NOTES

1. A more precise version of this reaction has been reported (Kaushik, S. M.; Yuan, Zhi; Noyes, Richard M. *J. Chem. Educ.* 1986, 63 (1), 76-80.). The more precise version uses 49.0 g of sodium perchlorate and calls for 11.2 mL of solution A and 13.1 mL of solution B.
2. Bursts of gas begin in 10-15 s and increase in intensity.
3. The less accuracy used in the preparation of this reaction, the more likely some gas will be produced continuously.
4. Some stirring or agitation is required if no gas is produced continuously.
5. The pH should be around 3.2 and vary up to 4.0. Acid concentration is an important factor.
6. Solutions can be stored for several weeks.
7. The perchlorate seems to have a "salt" effect on the rate of the reaction. Experiment with the concentrations.
8. Good mixing at the beginning of the reaction is important.
9. The explanations of the mechanism include the suggestion that as the solution becomes very supersaturated and the bubbles form on nucleation spots, the gas is evolved. A short amount of time is required for the concentration to again build up.
10. You may prefer to use sodium perchlorate (up to 50 g).

QUESTIONS FOR STUDENTS

1. How long does the gas evolution last?
2. What is the time interval between evolutions?
3. Discuss what is happening to the pressure in the test tube as the reaction oscillates.
4. As the reaction proceeds, how many different gases do you observe?
5. Find out what is meant by *nucleation*.

Reactions of Some Elements



27. Producing Hydrogen Gas from Calcium Metal

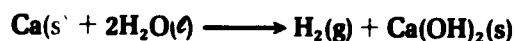
A small piece of calcium metal is placed in a beaker of water. As hydrogen is produced, a funnel is inverted over the metal to direct the gas into a test tube, where the gas is collected by the displacement of water.

Procedure

1. Fill a 400-mL beaker about three-fourths full of water.
2. Add 2 drops of phenolphthalein solution.
3. Prepare one or two test tubes for hydrogen collection. Fill them with water and invert the full tubes in the beaker of water.
4. Select a glass funnel small enough to be completely under water when inverted in the beaker. Add more water if necessary.
5. Drop a small piece of calcium metal (pea size) into the center of the beaker.
6. Immediately cover the area where the metal is with an inverted funnel.
7. As the hydrogen gas is produced and escapes through the stem of the funnel, place one of the test tubes over the stem of the funnel to collect the gas by displacement of water.
8. Collect a second tube of hydrogen. Keep the mouth of the tubes directed downward.
9. Test for hydrogen by inserting a glowing wooden splint into the mouth of the inverted test tubes.
10. Note the pink color in the solution and the precipitate.

Reactions

1. The calcium metal reacts with the water to produce hydrogen gas.



2. The hydrogen burns to produce water. When mixed with oxygen, hydrogen ignites and produces a "bark". This reaction is the positive test for hydrogen gas.



3. The basic calcium hydroxide causes the phenolphthalein to turn pink.

Teaching Tips

NOTES

1. The calcium hydroxide produced in the reaction is an insoluble white solid. It settles to the bottom of the beaker as a white precipitate.
2. Calcium is safe to use to prepare hydrogen. The alkali metals (sodium, lithium, and potassium) react so readily that the hydrogen ignites from the heat generated.
3. Aluminum, zinc, and magnesium will also liberate hydrogen if their oxide coatings are removed and the water is boiled. Finely divided zinc may displace hydrogen from cold water.

QUESTIONS FOR STUDENTS

- 1. Describe the reactions in this demonstration.**
- 2. List common metals according to their ability to produce hydrogen when placed in water.**
- 3. Describe the properties of hydrogen gas.**

28. Plastic Sulfur

Yellow sulfur is heated and forms an orange-red liquid. The liquid is poured into a beaker of water, and a flexible brown polymer of "plastic sulfur" is formed.

Procedure

1. Place a small amount of sulfur (to a height of about 1 in.) in a large test tube. Notice the properties of the sulfur.
2. Gently heat the test tube until the sulfur melts. Observe the formation of an orange liquid.
3. Continue to strongly heat the test tube. Notice the formation of a thick, dark red liquid.
4. Quickly pour the molten sulfur into a beaker of cold water. Observe the formation of a brown material, plastic sulfur.
5. Remove the rubbery material and note its properties including elasticity.

Reactions

1. Ordinary yellow sulfur, called *orthorhombic sulfur*, consists of S_8 molecules in a ring structure.
2. As the S_8 molecules are heated, the ring structure breaks and forms small 16-atom chains. This form of molten sulfur has an orange color.
3. Upon further heating (about 160°C), the 16-atom chains break and form long chains of sulfur atoms. This structure gives the molten sulfur a thick (viscous) appearance.
4. When the long chains of sulfur are suddenly cooled, the material keeps its long-chain structure, without the ordered arrangement of sulfur atoms found in the crystalline solid. This structure gives the plastic sulfur a "rubbery" character.

Teaching Tips

NOTES

1. Another form of sulfur, monoclinic, forms when liquid sulfur gradually cools. This form of sulfur has long, needlelike crystals.
2. The different types of sulfur are *allotropes*.
3. Try stretching the plastic sulfur. Test it for other rubbery characteristics.
4. Upon standing, the plastic sulfur will revert to the original orthorhombic form.
5. The melting point of yellow sulfur is 113°C .
6. Heating sulfur may produce some sulfur dioxide gas, which may be irritating. Do this demonstration in a hood or in a well-ventilated area.

QUESTIONS FOR STUDENTS

1. What happens, on a molecular level, as ordinary sulfur is heated?
2. Why does the plastic sulfur have properties of a rubber?
3. Is plastic sulfur a *polymer*?
4. Why does the sulfur become thick (viscous) when it is heated?

29. Recycling Aluminum

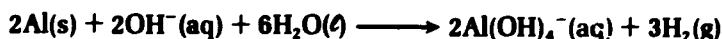
An aluminum can or foil is reacted with a base and then with sulfuric acid. Crystals that form are called *alum*.

Procedure

1. "Open" an aluminum soft drink can with a pair of scissors or use about 1 g of aluminum foil.
2. Scrape the paint and plastic coating from the can.
3. Cut the aluminum sample into bits (about 0.5 cm long)
4. Place the samples in 50 mL of 4 M KOH in a 250-mL beaker.
5. React until hydrogen is no longer evolved.
6. Filter and collect the filtrate in a 250-mL beaker.
7. Add 25–30 mL of 9 M sulfuric acid to the filtrate.
8. Place the mixture in an ice bath and observe the formation of large beautiful crystals.

Reactions

1. Aluminum reacts with potassium hydroxide to form soluble tetrahydroxoaluminate ion and hydrogen gas.



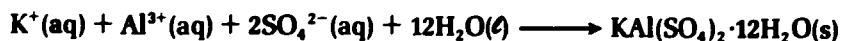
2. Addition of sulfuric acid forms the precipitate of aluminum hydroxide.



3. Additional acid neutralizes the hydroxide to form the clear solution of aluminum(III) ion and water.



4. The alum precipitates on cooling.



Solutions

1. The potassium hydroxide concentration is 4 M: Dissolve 112 g of KOH in 500 mL of solution.
2. The sulfuric acid concentration is 9 M: Add concentrated sulfuric acid to 50 mL of water until the total volume is 100 mL. BE CAREFUL with the concentrated acid and with the diluted acid. The solution will be quite hot. Always add acid to water when diluting.

Teaching Tips

NOTES

1. This series of reactions may take a little time but that time can be used to develop balanced equations, to discuss the importance of recycling aluminum, and, if the sample is preweighed, to predict the mass of the resulting alum crystals
2. The average predicted longevity of an aluminum can alongside the roadway is 100 years.
3. The energy requirement to prepare an aluminum can from recycling is only 5% of the energy required to produce the can from bauxite ore.
4. Over 90% of bauxite must be imported.
5. See demonstration 17 for a procedure for collecting the hydrogen gas in a balloon.
6. Alums are ionic compounds that crystallize from solutions containing sulfate, a trivalent cation (like Al^{3+} or Cr^{3+}), and a monovalent cation (like Na^+ or K^+).
7. The crystals can be grown to large sizes.
8. Alum is widely used as a coagulant in water purification, in the paper industry, and in making pickles.
9. The tetrahydroxoaluminate ion is also called the aluminate ion.

QUESTIONS FOR STUDENTS

1. Describe the crystals of alum.
2. What must happen to the aluminum in the alum to produce the metal?
3. Describe your observations and the reactions as the acid is added to the solution.
4. How could you "grow" a large crystal of alum?

30. Making Sodium Chloride from Sodium and Chlorine

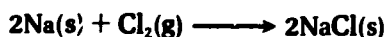
A deflagrating spoon containing molten sodium metal is lowered into a container of chlorine gas. A vigorous exothermic reaction occurs, and sodium chloride (table salt) is produced.

Procedure

1. Prepare the chlorine gas as follows. **THIS PROCEDURE SHOULD BE DONE IN A HOOD. CHLORINE GAS IS TOXIC AND CORROSIVE.**
 - a. Place about 3 tsp of calcium hypochlorite, $\text{Ca}(\text{OCl})_2$, in a sidearm flask with a stopper to fit. Add enough water to make a slurry. Attach a piece of rubber tubing to the sidearm flask and direct the other end to the bottom of a large round-bottomed flask.
 - b. When you are ready to generate and collect chlorine gas, add about 25-30 mL of 6 M hydrochloric acid (HCl), immediately stopper the flask, and collect the gas in the round-bottomed flask by upward displacement of air until the flask is full.
 - c. Cover the flask with a glass plate, remove the rubber stopper from the reaction vessel, and slowly fill the vessel with water to stop the chlorine-generating reaction.
 - d. When the contents of the reaction vessel are cool, flush them down the drain.
2. Place a pea-size amount of clean sodium in a deflagrating spoon and gently heat until the sodium melts.
3. Remove the cover from the container of chlorine and slowly lower the molten sodium into the gas.
4. Prepare for a bright and vigorous reaction.

Reaction

Sodium metal reacts with chlorine gas to produce the solid, sodium chloride.



Teaching Tips

NOTES

1. This reaction produces an intense light and a considerable amount of heat.
2. You could, of course, place the molten sodium in the flask generating the chlorine gas. However, this procedure is not recommended because of the possibility that molten metal could fall into the liquid.
3. The melting point of sodium metal is 97.8 °C.
4. The sodium chloride may have a yellow tint because of the formation of iron chloride from the metal spoon.
5. You should not allow students to taste the product to determine if it is salt. Test the product by adding a few drops of silver nitrate solution to a solution of the combustion product. A white precipitate of silver chloride shows the presence of chloride from the sodium chloride.

QUESTIONS FOR STUDENTS

1. Describe the reactants and products in this reaction.
2. What kinds of energy are involved in this reaction?
3. How can you determine if sodium chloride was actually produced?

31. Burning Magnesium in Carbon Dioxide

A burning magnesium ribbon, when placed into a flask containing dry ice, continues to react vigorously with the carbon dioxide to form black flakes of carbon.

Procedure

1. Place several small pieces of dry ice in a medium-sized beaker. Allow the beaker to fill with carbon dioxide gas.
2. Holding a 5-cm piece of magnesium ribbon with crucible tongs, light one end of the ribbon and insert it into the beaker.
3. Continue to hold the burning strip inside the beaker until the reaction ceases.

Reaction

The heat from the burning magnesium is sufficient to break down the carbon dioxide; magnesium oxide and carbon are produced.



Teaching Tips

NOTES

1. If you do not have dry ice, generate the carbon dioxide by reacting sodium bicarbonate and vinegar in a sidearm flask. Place a piece of tubing on the side arm and direct the flow of gas from the tube into a beaker.
2. Try not to touch the side of the beaker with the burning ribbon.
3. This demonstration is a good way to introduce oxidation-reduction reactions. The magnesium is oxidized and the carbon is reduced.
4. CAUTION students about the intense brightness of the burning magnesium ribbon.

QUESTIONS FOR STUDENTS

1. Discuss the kinds of energy evolved during the course of this reaction. (Light and heat.)
2. Write the half-reactions for the oxidation and reduction processes in this reaction.
3. Describe the bond-breaking and bond-forming processes in this reaction.

32. The Glowing Test Tube

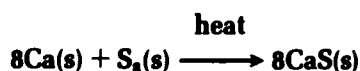
Small amounts of calcium metal and sulfur are placed in a test tube. The test tube is heated strongly over a Bunsen burner. After a few seconds, the tube glows brightly as the elements combine.

Procedure

1. Place a small piece of calcium metal (about the size of a match head) in a small test tube.
2. Add enough sulfur to cover the calcium.
3. Heat the test tube, gently at first and then strongly, with a burner.
4. Observe the bright glow when the elements react.

Reaction

Calcium metal combines with sulfur to produce calcium sulfide.



Teaching Tips

NOTES

1. A large amount of heat is produced in this reaction—enough to partially melt the tube. You must use a test tube that can be discarded after the demonstration.
2. A similar reaction occurs between iron and sulfur. Mix equal parts of iron filings and sulfur. Place about 1 in. of the mixture in a test tube and heat strongly.
3. This demonstration is a good example of a reaction that produces heat and light.
4. This demonstration is especially effective in a darkened room.
5. Ordinary sulfur is S_8 , a ring structure. Sulfur atoms exist only in the vapor phase above 2000 °C. However, sulfur is commonly represented as S in simple chemical reactions.
6. This reaction requires intense heat. Use the hottest Bunsen flame, or a Meker or Fisher burner.

QUESTIONS FOR STUDENTS

1. Why is heat required to initiate this exothermic reaction?
2. Why is the sulfur shown in the reaction written as S_8 ?
3. Which element lost electrons, and which gained electrons? (Calcium lost electrons, and sulfur gained electrons.)
4. How are the properties of the cooled calcium sulfide different from those of the reactants?

33. Halogens Compete for Electrons

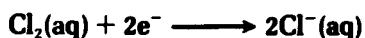
Water solutions change from colorless to yellow and brown and then to more surprising colors in another solvent when various combinations of halogen solutions are mixed.

Procedure

1. Prepare nine large test tubes, three each of 10 mL of solutions of KCl, KBr, and KI.
2. To the three test tubes of KCl add 10 mL of chlorine water to the first, bromine water to the second, and iodine water to the third. Set the test tubes aside.
3. Repeat procedure 2 with each test tube of KBr solution.
4. Repeat with each test tube of KI solution.
5. Shake each test tube and make observations.
6. Test the result of adding a little trichlorotrifluoroethane (TTE) to each of the stock halogens and halogen salt solutions. Stopper and shake the test tubes.
7. Note the appearance of molecular chlorine, bromine, and iodine in solution with TTE.
8. Add 10 mL of TTE to each of the nine test solutions from procedures 2, 3, and 4.
9. Account for the results you observe.

Reactions

1. The tendency for each halogen to gain and retain electrons is evident. Chlorine has the greatest tendency, then bromine, and then iodine. For example



2. Iodine has the greatest tendency to give up electrons (become oxidized).



Solutions

1. The salt solution concentrations are 0.2 M: KCl, 7.5 g/500 mL; KBr, 12 g/500 mL; and KI, 17 g/500 mL.
2. Chlorine water: 100 mL of commercial bleach and 75 mL of 2 M HCl.
3. Bromine water: 150 mL of KBr solution; 0.1 g of potassium bromate, KBrO_3 , and 25 mL of 2.0 M HCl.
4. Iodine water: 150 mL of water; a few grains of potassium iodate, KIO_3 ; 2 droppersful of KI solution; and 25 mL of 2.0 M HCl.
5. TTE, trichlorotrifluoroethane, is less toxic than the other nonpolar solvents.
6. KIO_3 , potassium iodate, solution: 2.1 g/500 mL.
7. TTE, trichlorotrifluoroethane, is less toxic than the other nonpolar solvents.

Teaching Tips

NOTES

1. Stoppered, these test tubes make a nice display for a relatively long time.
2. Students can construct the reactivity series of the halogens and find electrode potentials for the reactions.
3. When finished with the test tubes containing TTE, pour the solutions together and predict the final result. The water can be removed from the top and the TTE and halogens evaporated in the hood on paper towels.
4. All the molecular halogens are toxic and should not be breathed or handled.

QUESTIONS FOR STUDENTS

1. Which halogen is the strongest oxidizing agent—takes electrons from other atoms? (Chlorine.)
2. Which halogen is most easily oxidized—gives up its electrons? (Iodine.)
3. Are the molecular halogens more soluble in water or TTE?
4. Predict what will happen if a solution of potassium iodide is added to a solution of bromine water.
5. What are the colors of the halogens in water solution and in TTE?

34. Separating Metallic Iron from Cereal

A slurry is made by mixing a bowl of iron-fortified cereal with water. A magnet is passed through the slurry, and iron particles adhere to the magnet and can be removed from the cereal.

Procedure

1. Select a sample of cereal that is iron-fortified, as indicated on the box label.
2. Add water and stir until the cereal becomes soggy.
3. Place a white magnetic stirring bar in the cereal and stir with a magnetic stirrer.
4. After 15-20 min, remove the magnetic stirring bar and note the slivers of iron covering the magnet.

Reaction

The slivers of iron are physically removed by the magnet.

Materials

Cereals with a label stating "reduced iron" or "100% iron" are most likely to have the iron filings. Total works well.

Teaching Tips

NOTES

1. Humans cannot absorb metallic iron. However, the metallic iron reacts with hydrochloric acid in the stomach to produce the oxidized form, Fe^{2+} . Fe^{2+} is the more useful state of iron. Fe^{3+} often forms from Fe^{2+} if it is exposed to the air. This method of fortifying cereal protects against the oxidation of iron and is approved by the Food and Drug Administration.
2. You may want to keep the stirrer going for 30 min to pick up the maximum amount of iron.
3. Some chemicals found naturally in vegetables complex with iron and prevent its absorption—oxalate and tartrate are examples.
4. You may want to determine the amount of iron quantitatively. Dissolve the iron in 1 M HCl. Oxidize the iron further with 3% H_2O_2 and determine the amount of iron spectrophotometrically as the red complex, FeSCN^{2+} .
5. Use white, Teflon-coated magnets for best results.

QUESTIONS FOR STUDENTS

1. How is stomach acid, 0.1 M HCl, helpful in making iron suitable for absorption?
2. How is iron used in the body? Is it used as iron(II) or as iron(III) or both? Consult reference books in bioinorganic chemistry.
3. Design some investigations of the reactions of iron with acid and other reagents.

35. Floating Pennies

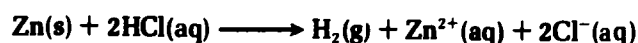
Copper-coated pennies (those made since 1983) are placed in a beaker containing hydrochloric acid. Within a few hours, the pennies mysteriously float to the top of the liquid.

Procedure

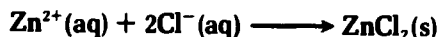
1. Place about 20 mL of hydrochloric acid in a 100-mL beaker.
2. Using a triangular file, make several small notches on opposite edges of a penny (you must use a "new" penny—one made since 1983. These pennies have a zinc core).
3. Place the penny in the acid.
4. Leave the reaction undisturbed overnight. Observe that the penny floats to the top of the liquid.
5. Remove the penny, carefully wash it, and examine it.

Reaction

1. The zinc in the core of the penny reacts with the hydrochloric acid to produce hydrogen gas and zinc chloride, which dissolves.



2. If the water is allowed to evaporate after the reaction, the solid zinc chloride residue will remain.



Solution

The hydrochloric acid concentration is 6 M: See Appendix 2.

Teaching Tips

NOTES

1. Be sure that the notches expose the zinc core. The amount of zinc exposed will affect the reaction rate.
2. If you use more pennies, use more acid. The zinc in one penny will require about 20 mL of acid.
3. Old pennies average 3.1 g each; new pennies (1983 and after) average 2.6 g. Old pennies contain 94% copper, new pennies contain 2.4% copper.
4. New quarters contain a copper core. Try reacting a new quarter with dilute nitric acid for the same effect.
5. This demonstration is a nice way to introduce the reaction of metals with acids.
6. Hydrogen gas trapped in the penny causes the penny to float.

QUESTIONS FOR STUDENTS

1. Explain the reaction that causes the pennies to "float".
2. Why does the copper not react with the acid?
3. What happens to the zinc metal?

36. Slow Copper Diffusion

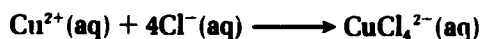
A test tube of blue crystals of copper sulfate is topped by white sodium chloride crystals and then a layer of water in which two nails rest. Over a period of several days, the salt layer becomes green and metallic copper appears on the nails.

Procedure

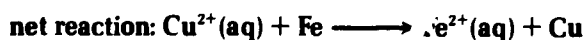
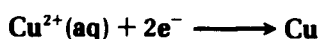
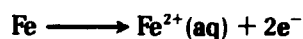
1. Place 2-3 cm of copper sulfate crystals in a test tube.
2. Place a paper circle on top of the crystals. (Cut circles from filter paper.)
3. Add sodium chloride to twice the thickness of the copper sulfate.
4. Place a paper circle on top of the salt.
5. Add water to the top of the test tube.
6. Place two iron nails in the water.
7. Observe over several days.

Reactions

1. Copper ions surrounded by water molecules, as in copper sulfate pentahydrate, are a blue color. As the copper ions diffuse into the sodium chloride layer, they complex with the chloride and form green CuCl_4^{2-} .



2. When copper ions contact metallic iron, the transfer of electrons from iron to copper occurs spontaneously.



Materials

1. Use solid copper sulfate pentahydrate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.
2. Use solid sodium chloride, NaCl .
3. Use noncoated iron nails.

Teaching Tips

NOTES

1. Challenge students to identify the reactions they are observing and to write balanced equations.
2. Set this reaction up early in the year so that students become accustomed to thinking about the difference between ions and metallic or elemental forms of substances.
3. It is useful to allow students to puzzle over their observations for several days.
4. Copper that is reduced by substances in nature is called *native copper*.
5. See Demonstration 39.

QUESTIONS FOR STUDENTS

1. What are the colors of copper ions in solution? Why are there different colors?
2. As the metallic copper forms, what happens to the iron?
3. How is copper metal produced from compounds in ore? Use the library or an encyclopedia.

37. Liquid Iodine

Solid iodine is placed in a test tube and stoppered with a one-hole stopper containing a long piece of glass tubing. When the tube is placed in an oil bath and heated to 145 °C, the iodine melts. Iodine vapors collect in the glass tubing.

Procedure

1. Place approximately 7 g of iodine in a 16 × 150 mm Pyrex test tube. The test tube is fitted with a one-hole stopper with a piece of glass tubing extending about 7 in. above the test tube.
2. Place the test tube in an oil bath and heat to about 145 °C. The iodine will melt at 113.5 °C.
3. Turn the heat off when the temperature reaches 145 °C. You want to avoid the boiling point of iodine (184.3 °C).

Reaction



Teaching Tips

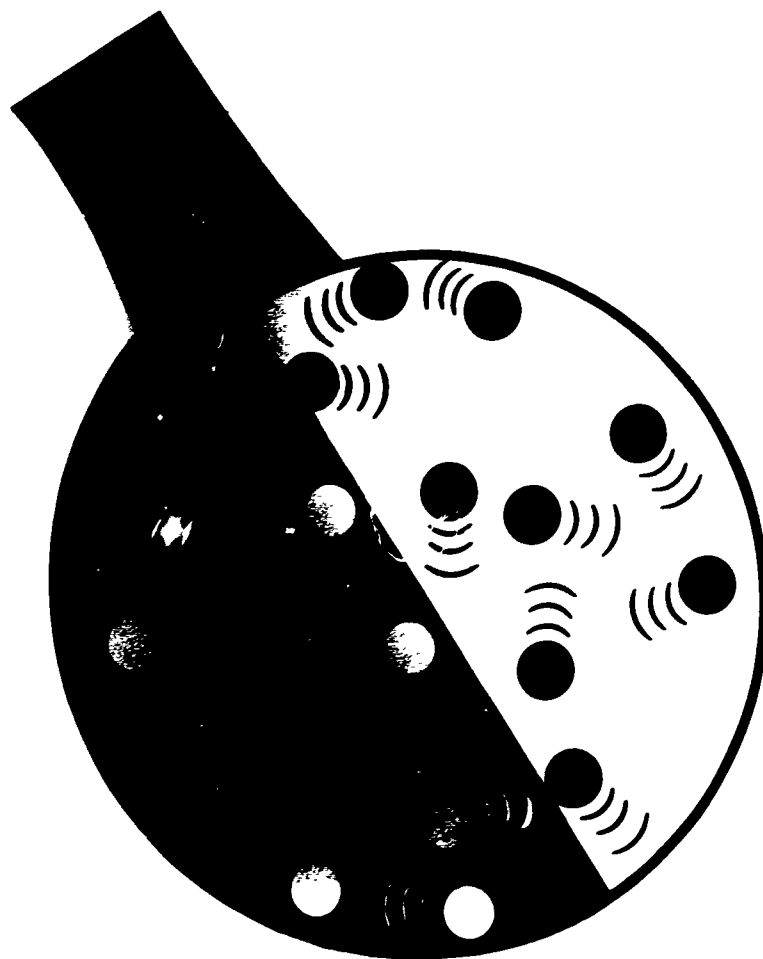
NOTES

1. Cottonseed oil works well for the oil bath.
2. As stated previously, the glass tubing allows for the collection of any iodine vapors that are produced during the experiment.
3. With the proper precautions, it is quite interesting to watch what happens when the liquid iodine is poured out into a Petri dish. You can observe the growth of iodine crystals. The solidified iodine is silvery and shiny, and when the Petri dish is covered, iodine vapors will be given off from the iodine. The obvious color of iodine vapors is quite evident. (But solid iodine does not look like it would have purple vapors!)

QUESTIONS FOR STUDENTS

1. What are the colors of solid, liquid, and gaseous iodine?
2. What do you observe as the melting point of iodine?
3. What happens in the test tube as the iodine melts?
4. Why was an oil bath used and not water?
5. Describe what happens when liquid iodine is poured into a Petri dish.

Transition Metals and Complex Ions



38. Copper Sulfate: Blue to White

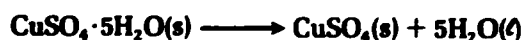
Blue crystals, when heated, turn white. Water is deposited on the beaker walls. Mass measurements can be taken on a top-loading balance.

Procedure

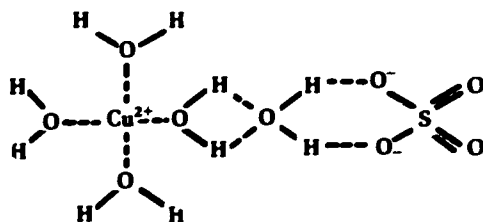
1. Place a few grams of fine crystals or powder of blue copper sulfate in a 250-mL beaker.
2. Put a watch glass on top.
3. Heat at the medium-low setting on a hot plate or over a low flame.
4. Observe the formation of water droplets on the watch glass and the inner walls of the beaker.
5. Remove the watch glass and continue heating until the crystals are white and no water remains.
6. Add a few drops of ethanol to one side of the anhydrous sample.
7. Add a few drops of water to the other side.
8. Feel the heat evolved as the hydration of copper sulfate occurs.

Reaction

Removing water molecules from the crystal is endothermic. The reverse reaction is exothermic. The water molecules are tightly held by the copper and sulfate.



The structure of the hydrated form is given as



Materials

1. Copper sulfate pentahydrate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.
2. A small amount of ethanol or other nonaqueous liquid.

Teaching Tips

NOTES

1. Fine crystals work best. The larger crystals take too long to lose water of hydration.

2. This demonstration is a good chance to make use of the electronic top-loading balance in a lecture experiment.
 - a. Weigh a 250-mL beaker.
 - b. Add 2.5 g of copper sulfate.
 - c. Record the mass.
 - d. Heat until the crystals are white.
 - e. Reweigh.
 - f. Have students calculate the mass and moles of water driven off.
 - g. Subtract the mass of the beaker and have students find the mass and moles of anhydrous copper sulfate remaining.
 - h. Find the ratio of moles of water to moles of copper sulfate. The ratio should be close to 5:1.
3. Notice that only water molecules can form the structure of the copper sulfate crystal that produces blue crystals. The addition of an unknown liquid to anhydrous copper sulfate is a test for water.

QUESTIONS FOR STUDENTS

1. How do you know blue copper sulfate is hydrated?
2. How do you know the hydrated form is more stable? (Heat evolves when it is formed.)
3. What is the most obvious difference between the hydrated form and the anhydrous form?
4. If you took measurements, what is the formula for the hydrated copper sulfate?

39. Green and Blue Copper Complexes

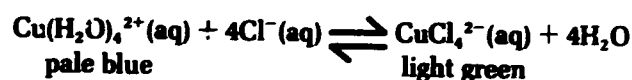
A blue solution is placed in a graduated cylinder. Hydrochloric acid is added to produce a green solution. When ammonia is added, a deep blue solution forms as a top layer in the cylinder.

Procedure

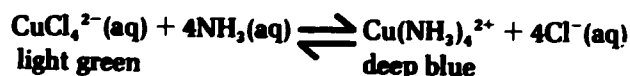
1. Fill a large graduate cylinder about one-fourth full with copper sulfate solution.
2. Carefully add hydrochloric acid until a pale green solution is formed.
3. Using a pipet, carefully add ammonium hydroxide until a deep blue color forms. With care, the deep blue solution will form a layer above the green solution.

Reactions

1. The green solution is due to the formation of the tetrachlorocopper(II) complex ion from the tetraaquocopper(II) formed when copper sulfate is in aqueous solution.



2. Aqueous ammonia forms the deep blue tetraamminecopper(II) complex with copper(II).



3. With care, the two solutions will not mix and will form separate layers, because the HCl is more dense than the aqueous ammonia layer.

Solutions

1. The copper sulfate concentration is 0.2 M: Dissolve 50 g of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in 1 L of solution.
2. Concentrated hydrochloric acid (CAREFUL).
3. Concentrated ammonium hydroxide.

Teaching Tips

NOTES

1. You can do this demonstration in separate beakers if you wish to use less concentrated acid and base.
2. The tetraaquocopper(II) and tetraamminecopper(II) complexes have a square-planar structure. The tetrachlorocopper(II) complex has the structure of a tetrahedron.
3. CAUTION: A considerable amount of heat is produced when the ammonium hydroxide is added to the acid.
4. You can expect to use about the same volume of acid and ammonium hydroxide to produce the best colors.

QUESTIONS FOR STUDENTS

1. Explain what effect the HCl has on the equilibrium expressed by the first reaction and what effect aqueous ammonia has on the equilibrium in the second reaction.
2. What is a *complex ion*?
3. What complex ion is responsible for the light blue color in a copper sulfate solution?
4. Why do the two layers form without mixing?

40. Changing Coordination Numbers: Nickel Complexes

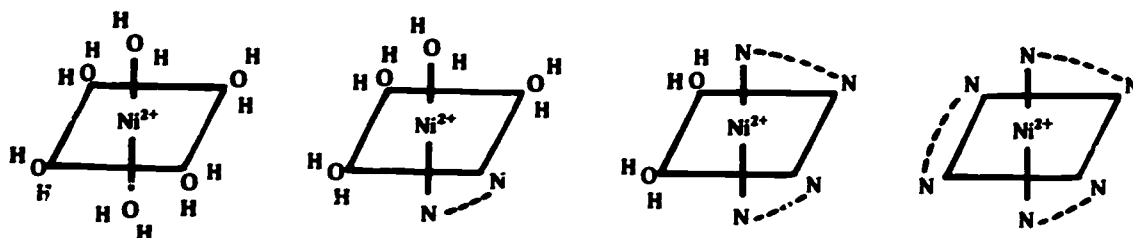
Four beakers containing a green solution of nickel(II) solution are displayed. When successively more drops of diaminoethane are added to the beakers, the solutions display colors from light blue to violet characteristic of various nickel complexes.

Procedure

1. Add 100 mL of nickel chloride solution to each of four 250-mL beakers.
2. Add 1 drop of diaminoethane (ethylenediamine) to the second beaker. Note the appearance of a light blue color.
3. Add 3-4 drops of diaminoethane to the third beaker. If a royal blue color does not appear, add a few more drops until it does.
4. Add enough diaminoethane to the fourth beaker to change the color to violet.

Reactions

1. The first beaker contains a green solution of hydrated Ni^{2+} ions, $\text{Ni}(\text{H}_2\text{O})_6^{2+}$.
2. The second beaker contains a light blue solution of the complex ion in which two water molecules have been replaced by one ethylenediamine (en), $\text{Ni}(\text{H}_2\text{O})_4\text{en}^{2+}$.
3. The third beaker contains the royal blue color characteristic of the nickel complex consisting of two water molecules and two en's, $\text{Ni}(\text{H}_2\text{O})_2(\text{en})_2^{2+}$.
4. The fourth beaker has no water in its complex ion structure but contains three en's, $\text{Ni}(\text{en})_3^{2+}$.
5. Structures of the four complex ions are shown. Note that the bidentate ligand ethylenediamine (en) is shown as N—N.



Solutions

1. The concentration of nickel chloride is not important in this demonstration. Add enough nickel chloride to 500 mL of distilled water to produce a green color.
2. Any other Ni^{2+} compound will work just as well.

Teaching Tips

NOTES

1. Ethylenediamine, or 1,2-diaminoethane, has the formula $\text{H}_2\text{NCH}_2\text{CH}_2\text{NH}_2$. Because this compound has two donor nitrogens, it is called a bidentate ligand.
2. After use, these solutions can be flushed down the drain with copious amounts of water.
3. You might want to ask students to predict the colors of these complex ions. For example, place 1 drop in beaker 2 and a much larger amount in beaker 4 and ask students to predict the intermediate color in beaker 3.
4. Bidentate ligands are often called *chelates*, from the Greek word meaning *claw*. Notice in the structure that the ligand "grasps" the nickel ion between two donor nitrogens.

QUESTIONS FOR STUDENTS

1. What is a *complex ion*?
2. Why is ethylenediamine (en) called a *chelating agent*?
3. How can you explain the different colors in the different beakers?
4. Explain the geometrical shape of these complexes.

41. Colorful Complex Ions in Ammonia

In an arrangement of beakers, solutions of copper, cobalt, and nickel ions are treated with ammonia to produce the colors characteristic of their ammine complexes.

Procedure

1. Arrange six beakers into three pairs.
2. To the first pair, add 100 mL of copper nitrate solution.
3. To the second pair, add 100 mL of cobalt nitrate solution.
4. To the third pair, add 100 mL of nickel nitrate solution.
5. Use one member of each pair as the color standard and treat the others to form the colored complexes by adding ammonia until the colored complexes (deep blue, red, and violet, respectively) are formed.
6. Compare the colors of the ions in solution with that of the ammine complex.

Reactions

The reactions that cause the color changes are replacements of the water molecules clustered about the metal ions by ammonia molecules. The metal ions form stronger bonds with ammonia than they do with water.

1. $\text{Cu}(\text{H}_2\text{O})_6^{2+}(\text{aq}) + 4\text{NH}_3(\text{aq}) \longrightarrow \text{Cu}(\text{NH}_3)_4^{2+}(\text{aq}) + 4\text{H}_2\text{O}$
2. $\text{Co}(\text{H}_2\text{O})_6^{2+}(\text{aq}) + 6\text{NH}_3(\text{aq}) \longrightarrow \text{Co}(\text{NH}_3)_6^{2+}(\text{aq}) + 6\text{H}_2\text{O}$
3. $\text{Ni}(\text{H}_2\text{O})_6^{2+}(\text{aq}) + 6\text{NH}_3(\text{aq}) \longrightarrow \text{Ni}(\text{NH}_3)_6^{2+}(\text{aq}) + 6\text{H}_2\text{O}$

Solutions

1. The concentrations of the solutions are unimportant. You may use any of the soluble salts of cobalt(II) (pink), copper(II) (light blue), and nickel(II) (light green).
2. Any concentration of ammonia may be used; however, a smaller volume of concentrated ammonia would be required.

Teaching Tips

NOTES

1. The correct name for the cobalt complex in reaction 2 is hexaamminecobalt(II). It has an octahedral structure with the cobalt ion in the middle and NH_3 at each point in the structure. The nickel complex is hexaamminenickel(II), also with an octahedral structure. The copper complex, tetraamminecopper(II), has a square-planar structure with a copper ion in the center of the square and ammonia at each of the four corners.
2. *Ammine* is the term used to designate a complex ion formed with ammonia molecules.
3. You may dispose of these solutions by flushing them down the drain.

QUESTIONS FOR STUDENTS

1. What is a *complex ion*?
2. Compare the colors of the aqueous complexes with those of the ammonia complexes.
3. Show the geometrical shape of each complex ion.

42. The Magic Handkerchief

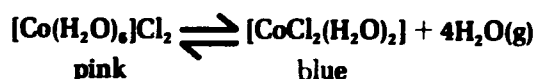
A blue handkerchief is crumpled and cupped in the hands. Breath is blown through the handkerchief for a minute whereupon the handkerchief turns white.

Procedure

1. Prepare the "magic handkerchief" by soaking it in a solution of cobalt(II) chloride and allowing it to dry.
2. If the dried handkerchief is not blue, warm it gently with a hair dryer or in an oven until it turns blue.
3. Crumple the handkerchief to increase its surface area and cup it into your hands.
4. Blow your breath into the handkerchief until the handkerchief turns from blue to pale pink or white.

Reaction

When heated, the pink cobalt complex turns blue because water is driven off. When moisture from the breath reacts with the blue complex, the handkerchief turns pink again.



Solution

The cobalt(II) chloride concentration is 2 M: Dissolve 26 g of cobalt(II) chloride in 100 mL of solution.

Teaching Tips

NOTES

1. The pink cobalt complex, hexa-aquacobalt(II) chloride, has an octahedral geometry. The blue cobalt complex, diaquodichlorocobalt(II), is tetrahedral.
2. The cobalt chloride solution poses no hazard and may be disposed of by flushing down the sink.
3. You can, of course, use tissues or white paper napkins instead of a handkerchief. Prepare several and store in a tightly sealed jar.

QUESTIONS FOR STUDENTS

1. What happens chemically when the blue handkerchief turns pink?
2. What does the breath have to do with the reaction?
3. What practical applications can you think of for this reaction? (Moisture detection.)

43. Chromate Dyes

A cotton handkerchief is folded and dipped into a colorless solution. The handkerchief is then dipped into a light orange solution and becomes bright YELLOW. The handkerchief is then dipped into a boiling solution and becomes bright ORANGE.

Procedure

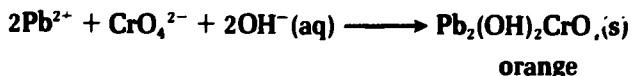
1. Place 100 mL of solution A in a large beaker.
2. Place 100 mL of solution B in another large beaker.
3. Place 100 mL of solution C in another beaker, and heat the solution until it is nearly boiling.
4. Fold a cotton handkerchief lengthwise several times so that it can be held and dipped into the solutions.
5. Dip the handkerchief into solution A. Note that no color is produced. Blot the wet handkerchief with paper towels.
6. Dip the wet end into solution B. Note the immediate formation of a YELLOW color.
7. Blot the yellow handkerchief, and dip it into the hot solution C. Note the immediate formation of an ORANGE color.

Reactions

1. Dipping the handkerchief into the first beaker saturates it with lead nitrate.
2. Dipping the handkerchief into the second beaker causes a reaction between the lead nitrate and potassium dichromate to form the YELLOW pigment, lead chromate.



3. Dipping the yellow handkerchief into the third beaker of calcium hydroxide causes the formation of lead chromate (basic), an ORANGE pigment.



Solutions

1. Solution A: lead nitrate solution. Dissolve 2 g of $\text{Pb}(\text{NO}_3)_2$ in 100 mL of water.
2. Solution B: potassium dichromate solution. Dissolve 1 g of $\text{K}_2\text{Cr}_2\text{O}_7$ in 100 mL of water.
3. Solution C: calcium hydroxide solution. Dissolve 5 g of $\text{Ca}(\text{OH})_2$ in 100 mL of water.

Teaching Tips

NOTES

1. Use a stirring rod to push the handkerchief into the solutions.
2. Be sure to remove excess solution after each dipping by blotting the handkerchief.

3. Allow the handkerchief to dry and try to remove the color.
4. Lead dyes are no longer used because of the danger of lead poisoning.
5. The PbCrO_4 forms from the basic solution of dichromate.
6. The orange product in the final step is a mixture of yellow PbCrO_4 (known as chrome yellow) and red $\text{Pb}_2(\text{OH})_2\text{CrO}_4$ (known as chrome red).
7. As with other heavy metal ions, the chromium solutions should be precipitated as sulfides and the solids discarded as described in Appendix 4.
8. You can use heavy white paper napkins also.

QUESTIONS FOR STUDENTS

1. Describe the colors of the solutions used in these reactions and the colors of the products that form on the handkerchief.
2. How is each colored product related to the solutions used?
3. Write an equation for each reaction forming a colored product on the handkerchief.

44. Appearing and Disappearing Silver

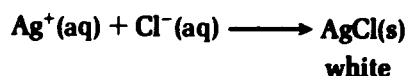
Silver ions show some interesting solubility chemistry. With added chloride, a cascade of precipitate forms that can be dissolved in excess of ammonia. When the ammonia is neutralized by nitric acid, the precipitate reappears. Adjustment of the pH of the solution causes a shift in equilibrium and the disappearance and reappearance of the precipitate of silver chloride.

Procedure

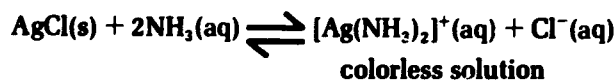
1. In a large test tube, place 10–25 mL of 0.1 M AgNO_3 .
2. Add 6 M HCl drop by drop until a dense white precipitate forms.
3. Add concentrated ammonia dropwise until the precipitate dissolves.
4. Add 6 M nitric acid dropwise until the precipitate re-forms.

Reactions

1. The white precipitate is insoluble silver chloride formed by the combination of silver ions with chloride ions.



2. The silver chloride precipitate dissolves in excess ammonia to form a stable complex ion.



3. The addition of excess HNO_3 to the clear solution lowers the concentration of NH_3 because NH_3 reacts with H^+ to form NH_4^+ . This reaction shifts the equilibrium to the left, and more white precipitate is formed.

Solutions

1. Silver nitrate: Dissolve 1.7 g of AgNO_3 in 100 mL of solution.
2. The concentration of HCl is 6 M: See Appendix 2.
3. Concentrated aqueous NH_3 .
4. The concentration of HNO_3 is 6 M: See Appendix 2.

Teaching Tips

NOTES

1. You may want to keep the silver chloride solid for future use. Flush any remaining solid. *Do not store the complex ion solution.*
2. The intermediate complex ion, $\text{Ag}(\text{NH}_3)^+$, is formed first when silver nitrate and ammonia are mixed.
3. This demonstration is useful to show solubility, equilibrium, and complex ion formation. The formation of the silver ammine complex is the step in classical qualitative analysis by which the silver is separated from mercury(I).
4. The complex ion, $[\text{Ag}(\text{NH}_3)_2]^+$, is the diamminesilver(I) ion.

QUESTIONS FOR STUDENTS

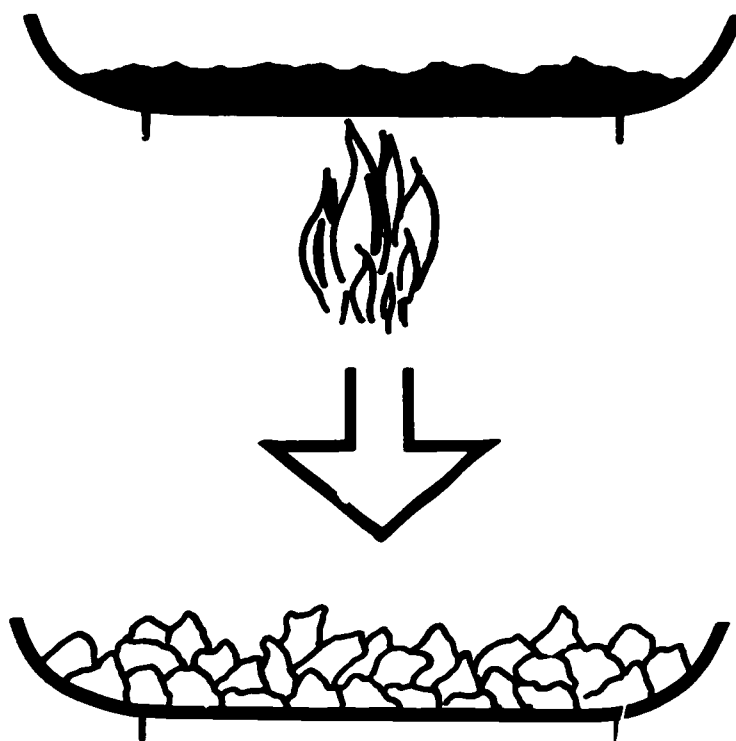
1. Describe, identify, and write an equation for the formation of the precipitate observed in the first reaction.
2. Why did the precipitate dissolve with excess ammonia?
3. What effect did adding nitric acid have on the equilibrium?
4. What is a *complex ion*?

2. The strong oxidizing effect of permanganate makes it useful for treatment of athlete's foot, for treatment of rattlesnake bite, and as an antidote for some poisons.
3. MnO_2 is the black material in dry cells. It is useful because of its strong oxidizing power and its high conductivity.
4. Chromium does not occur as the element in nature.

QUESTIONS FOR STUDENTS

1. Write the equations for the reactions.
2. Are the ions used in this demonstration only stable in basic or acidic solutions? Which in which?
3. Would another acid have worked just as well?
4. What was the role of the nitrate ion?
5. Could you have used something other than nitrate? What?

Chemical Bonding



46. Microcrystal Formation

Watching crystals form and grow is exciting evidence of the particle nature of matter.

Procedure

1. Place 1 tsp of sodium thiosulfate pentahydrate, phenyl salicylate, or sodium acetate trihydrate on a glass plate.
2. Melt the crystals on a warm hot plate or over a candle.
3. Place the plate on an overhead projector, drop on one "seed" of crystal, and watch the orderly crystallization.
4. Prepare a concentrated solution of glycine.
5. Place a pool of the glycine solution on a glass plate and observe crystals form as the solvent evaporates.

Reaction

All of these phase changes are physical changes in which the substance is melted or dissolved in water of crystallization and then solidifies or crystallizes from solution.

Materials and Solutions

1. Sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_5 \cdot 5\text{H}_2\text{O}$, is also called *hypo*.
2. Phenyl salicylate, $\text{C}_6\text{H}_4(\text{OH})\text{COOC}_6\text{H}_5$, is also called *salol*.
3. Sodium acetate, $\text{NaCH}_3\text{COO} \cdot 3\text{H}_2\text{O}$.
4. Glycine, $\text{CH}_3\text{NH}_2\text{COOH}$, is an amino acid. Add 1 g to 2 mL of warm water.

Teaching Tips

NOTES

1. Melting points of the suggested solids are 40 °C for sodium thiosulfate, 45 °C for phenyl salicylate, and 58 °C for sodium acetate.
2. Each of these crystal formations may be projected on a microprojector or viewed by students under a microscope.
3. If you can obtain photographic slide protectors made of glass, sandwich the melt between two of them and slip the slide into a slide projector for a spectacular view of crystallization on a screen.
4. Try projecting crystal formation of many types—saturated solutions of potassium nitrate, sodium sulfate, and aluminum or chromium alums. Solid benzoic acid melts at 122 °C, and solid resorcinol (1,3-dihydroxybenzene) melts at 110 °C. Both form nice crystals.
5. Borrow polarizing films from the physics teacher and experiment with them above and below the crystals. Polarizing filters are used for much of the impressive microphotography of crystals.

QUESTIONS FOR STUDENTS

1. Describe the "growth" of the crystals.
2. In each case, what is the chemical composition of the crystal?
3. What is the difference between solidifying from a melt and crystallizing from a saturated solution?

47. Solubility and Immiscibility

Two Petri dishes are placed on an overhead projector. Each Petri dish contains a pool of trichlorotrifluoroethane (TTE) in water. Iodine dissolves in the pool in one Petri dish but not in the surrounding water. Copper sulfate does not dissolve in the pool in the second dish but does in the surrounding water.

Procedure

1. Place two Petri dishes on a plastic sheet on an overhead projector.
2. Add enough water to each dish to cover the bottom.
3. In each dish, carefully add about 2 mL of TTE. Notice that it forms a pool.
4. Drop three to four iodine crystals in the pool and the same number in the water near the pool. Observe the solubility of iodine in TTE but not in water.
5. In the second dish, add a few crystals of copper sulfate to the pool and the same amount to the water near the pool. Observe the solubility of CuSO_4 in water but not in TTE.

Reactions

1. In the first Petri dish: Iodine, I_2 , is a nonpolar substance, so it dissolves in the nonpolar solvent TTE to produce a purple color. It does not dissolve in the polar water.
2. In the second Petri dish: Copper sulfate, CuSO_4 , is a polar (ionic) substance and dissolves in the polar water to produce a blue solution but not in the nonpolar TTE.

Teaching Tips

NOTES

1. TTE is the accepted substitute for carbon tetrachloride, CCl_4 . TTE is readily available from chemical supply companies.
2. The formula for TTE is $\text{CCl}_2\text{FCClF}_2$. It has a density of 1.6 g/mL.
3. You can, of course, do the same demonstration in graduated cylinders or test tubes. The Petri-dish method uses smaller amounts of chemicals.
4. This demonstration is a good way to illustrate the idea that "like dissolves like".

QUESTIONS FOR STUDENTS

1. Explain why each substance dissolves as it does.
2. Why do water and TTE not mix?
3. What would happen if you used potassium permanganate instead of copper sulfate in this demonstration? Try it!

48. Bending a Stream of Water

Mysterious powers seem to be at work as a stream of water from a buret is bent (attracted to a "charged" rubber rod). A stream of hexane is not bent.

Procedure

1. Fill two burets with two liquids—water in one and a nonpolar solvent like hexane in another.
2. "Charge" a rubber or plastic rod by rubbing it with a wool or fur piece.
3. Open the stopcock of the buret containing the water until a fine unbroken stream flows out.
4. Hold the charged rod near the stream and observe that the stream is attracted to the rod.
5. Repeat with the hexane. The stream is unaffected.

Reaction

The water molecule is polar, slightly negative at the oxygen atom and slightly positive at the hydrogens. The molecules are attracted to a charged object. The hexane molecule has no polarity.

Solutions

1. Hexane or another nonpolar liquid.
2. Water.

Teaching Tips

NOTES

1. You may want to demonstrate more than one nonpolar liquid. Try any of the hydrocarbons or trichlorotrifluoroethane.
2. The charged rod could be glass rubbed with silk.
3. In 1600, Dr. William Gilbert, a physician to Queen Elizabeth and the father of the modern science of electricity and magnetism, said that the rubber rod stroked with fur produces "negative electrification". The glass rod stroked with silk produces "positive electrification".
4. Polarity of water accounts for its liquid state at room temperature, its lower density as ice, its ability to dissolve ions, and its tendency to form ions, H^+ and OH^- .

QUESTIONS FOR STUDENTS

1. Explain how a stream of water is bent by a charged rod although a stream of hexane is not.
2. What properties of water are explained by its polarity?
3. What properties of hexane are explained by its nonpolar nature?
4. Name other polar and nonpolar molecules.

49. Waves in a Bottle

Blue liquid creates elegant wave action against a colorless liquid in a plastic soft drink bottle.

Procedure

1. Half fill a 1-L clear plastic bottle with distilled water.
2. Add 50 mL of ethyl alcohol.
3. Add charcoal lighter fluid to fill the bottle to the brim.
4. Add four drops of blue food coloring, one at a time.
5. Place the lid on the bottle and seal it with tape.
6. Turn the bottle on its side and rock it to produce waves.

Reactions

Water and food coloring solutions are polar. Charcoal lighter fluid is nonpolar and less dense than water.

Solutions

1. Water.
2. Food coloring.
3. Charcoal lighter fluid, a mixture of hydrocarbons such as pentane and hexane, is sometimes called *petroleum distillates*. Paint thinner may be used.
4. Ethyl alcohol.

Teaching Tips

NOTES

1. Try other combinations of liquids and other colors.
2. It is **IMPORTANT** that a plastic container be used because it will be flexible if the fluids exert some pressure on a warm day.
3. The lighter fluid is flammable.
4. If ethyl alcohol is not available, try other alcohols or acetone.
5. Duct tape works well for sealing the bottle.

QUESTIONS FOR STUDENTS

1. In which liquid did the food coloring dissolve?
2. Why did the food coloring dissolve in only one liquid?
3. Why do the liquids not mix?
4. Why is the blue layer always below the colorless layer?

50. The Nonpolar Disk Game

Ten paper disks that are blackened on one side line up with the black side down with high frequency when they are shaken in a flask with two immiscible liquids.

Procedure

1. Using a white index card, blacken one side with the lead of a pencil. Punch out the disks or cut them about 0.5 cm in diameter.
2. Place about 100 mL of trichlorotrifluoroethane (TTE) in a 250-mL flask.
3. Pour in an equal volume of water.
4. Place the 10 disks in the flask, stopper the flask, and shake the flask.
5. Observe the number of disks that are oriented in each direction as the layers form.

Reactions

The graphite that forms the pencil "lead" is nonpolar carbon, which associates with the nonpolar solvent (TTE). The cellulose that is exposed paper on the other side of the disk has polar regions that are attracted to the water, which is also polar.

Teaching Tips

NOTES

1. Apply a heavy coating of graphite. The paper soon becomes soaked with water.
2. Try using a nonpolar solvent such as charcoal lighter fluid or hexane, which is less dense than water so the graphite orients up instead of down.
3. Try adding a few drops of detergent to the flask.
4. The phenomenon observed is explained by polar and nonpolar interactions—those interactions also observed in the action of soap molecules, formations of emulsions, and the structure of biological membranes.

QUESTIONS FOR STUDENTS

1. Determine the frequency per shake of the positions of the disks. How many disks were turned "up" or "down" each time?
2. What do the graphite and the nonpolar solvent have in common?
3. What other tests could you perform to show the relationships between the solvents and the sides of the disks?

51. Alkanes versus Alkenes: Reaction of the Double Bond

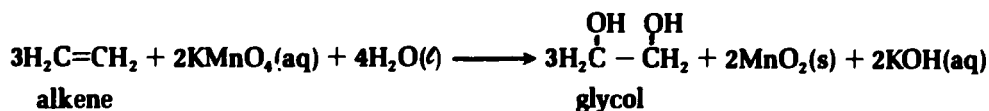
A clear liquid alkane (cyclohexane) and a clear liquid alkene (cyclohexene) are placed in separate beakers. Potassium permanganate is added dropwise to each beaker. The purple color characteristic of the permanganate persists in the alkane but fades in the alkene.

Procedure

1. Half fill a large beaker with cyclohexane.
2. Half fill a second beaker with cyclohexene.
3. Add potassium permanganate solution, dropwise, to the first beaker. Notice that the purple color of the permanganate persists in the liquid. Swirl the beaker.
4. Add potassium permanganate solution, dropwise, to the second beaker. Notice that the color fades and a brown precipitate forms. Swirl the beaker.

Reaction

Although this reaction is not specific for the double bond, cyclohexene with its one double bond does react, whereas cyclohexane with only single bonds does not react. The alkene is oxidized by permanganate to a glycol. The brown precipitate, MnO_2 , results from the reduction of KMnO_4 .



Solutions

1. The potassium permanganate concentration is 0.01 M: Dissolve 0.1 g of KMnO_4 in 100 mL of solution.
2. Any alkane or alkene will work. Cyclohexane and cyclohexene are common hydrocarbons and are safe to use.

Teaching Tips

NOTES

1. This demonstration projects well and can be done in Petri dishes on an overhead projector.
2. This reaction is called the *Baeyer test*. Alcohols and aldehydes also give a positive test, because they have easily oxidized groups.
3. A glycol is an alcohol with two -OH groups. A glycerol has three.
4. Although these hydrocarbons are generally safe, you should avoid inhaling their vapors.

QUESTIONS FOR STUDENTS

1. What is an *alkane*? What is an *alkene*?
2. Why is an *alkene* more reactive than an *alkane*?
3. Explain why the color does not persist in the *alkene* when permanganate is added.
4. What is oxidized in this reaction? What is reduced?

52. The Disappearing Coffee Cup

An expanded polystyrene coffee cup is placed in a Petri dish containing a solvent. The cup appears to "melt" and gets smaller and smaller as it sinks into the dish until it finally disappears.

Procedure

1. Add acetone to a Petri dish until it is about three-fourths full.
2. Place an empty polystyrene coffee cup in the center of the dish.
3. Observe.

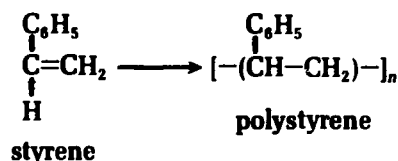
Reaction

Some solvents actually dissolve the polymer material; others merely break certain bonds that give the polystyrene cup its shape. In these solvents, like acetone, you can reclaim the glob of material after the cup disappears.

Teaching Tips

NOTES

1. You can use the common white coffee cup or polystyrene packing material.
2. Polystyrene is one of the simplest and more common polymers. It consists of repeating units of the styrene molecule.



3. Polymers like polystyrene are *thermoplastic*, which means that they soften when heated. In this softened state, they can be molded into cups, trays, and so forth.
4. This plastic is chemically stable and does not conduct heat well—making it ideally suited for use as insulating material.
5. In 1986, over 6 billion pounds of polystyrene was produced.
6. Reclaim and wash the sticky residue from the cup. The residue will eventually harden again.

QUESTIONS FOR STUDENTS

1. What is *polystyrene*?
2. What properties of this polymer make it suitable for making coffee cups? What else?
3. How does the polystyrene react in the solvent?

53. Hydrogen Bonding in Slime

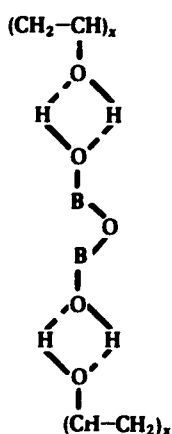
Two clear liquids are mixed. The result is a shimmery fluid polymer that can be shaped, that flows, and that will sheer if twisted quickly.

Procedure

1. Measure 50 mL of poly(vinyl alcohol) solution into a paper cup or small beaker and observe its properties.
2. Measure 7–8 mL of sodium tetraborate solution into another cup or beaker and observe its properties.
3. Choose a favorite color and add 1 drop of food coloring to the sodium tetraborate solution.
4. Pour sodium tetraborate into the poly(vinyl alcohol) solution while stirring vigorously with a wooden stick.
5. Examine the properties of the cross-linked polymer.

Reactions

1. Poly(vinyl alcohol) is a polymer derived from the monomer vinyl alcohol.
2. The borate forms a complex structure called tetraborate, $B_4O_5(OH)_4^{2-}$, that links the poly(vinyl alcohol) polymer strands together by hydrogen bonds.



Solutions

1. Poly(vinyl alcohol) is a 4% solution: Add 4 g to 100 mL of water. Heat to 70 °C until the solution clears. Do not boil.
2. Sodium borate is a 4% solution: Add 4 g of sodium borate, $Na_2B_4O_7 \cdot 10H_2O$, to 100 mL of water.

Teaching Tips

NOTES

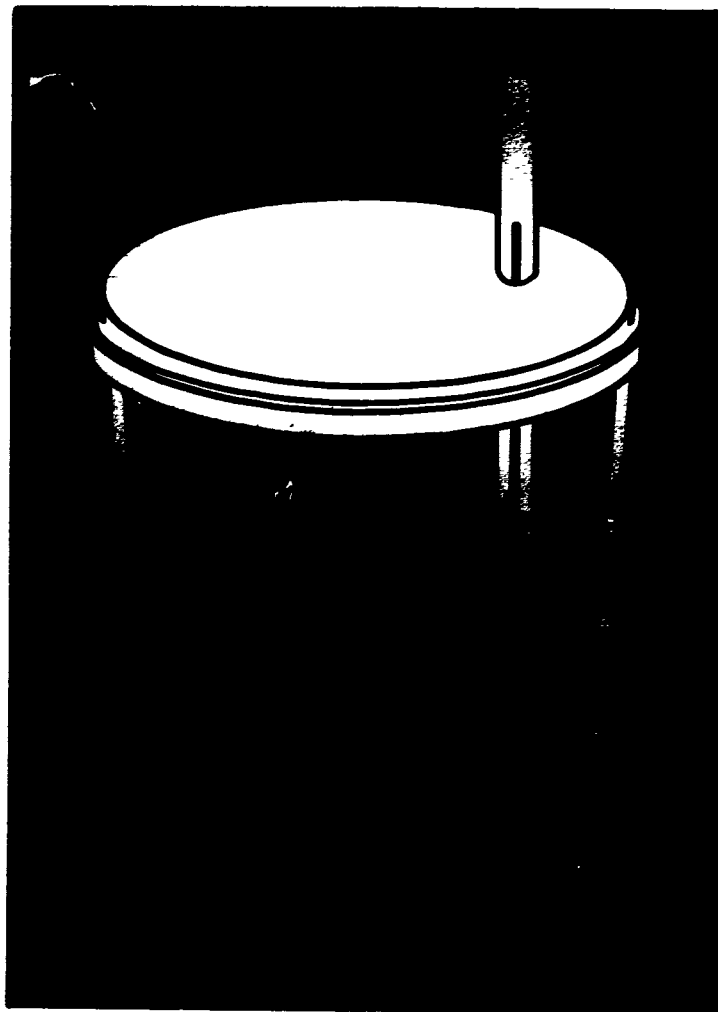
1. The poly(vinyl alcohol) is a fine dust. Avoid inhaling it.
2. Green is a popular food coloring for this polymer.

3. Experiment to see how far the polymer will flow from your hand.
4. It is interesting that the flowing is endothermic. The students will notice that it is cold.
5. If pressed onto a design written with magic marker, the slime will pick up the design.
6. The polymer is nontoxic.
7. Cross-linking can be compared with hydrogen bonding that links proteins into specific structures.
3. Sodium borate is also called sodium tetraborate.

QUESTIONS FOR STUDENTS

1. How did the properties of the poly(vinyl alcohol) change after it was cross-linked?
2. Describe the properties of the "slime".
3. Do you think the polymer is "slimy"? Suggest a better name.
4. What does it mean if the polymer is cold while it is flowing?

Energy Changes



54. A Chemical Hand Warmer

Several materials are placed in a plastic baggie. When the substances are mixed, a slow and steady evolution of heat results. This reaction is similar to that of the commercially available hand warmers.

Procedure

1. Place about 25 g of iron powder in a small plastic baggie.
2. Add 1 g sodium chloride.
3. Close the bag and shake to mix the chemicals.
4. Add about 1 Tbs of small vermiculite and mix again. The "hand warmer" is now ready to be activated.
5. When you are ready for the demonstration, add 5 mL of water to the bag and seal it with a metal twist-tie.
6. Squeeze and shake the bag to thoroughly mix the contents.
7. After a minute or so, a noticeable amount of heat will be produced.
8. Pass the hand warmer around the class.

Reactions

1. This demonstration is essentially an oxidation-reduction reaction. Iron is oxidized in an exothermic reaction.



2. The oxygen is reduced, picking up the electrons lost by the iron.
3. Sodium chloride provides electrolytes to catalyze the reaction.
4. Vermiculite helps insulate the hand warmer and contain the heat and distribute it in the bag.

Material

Iron powder: 100 mesh works well.

Teaching Tips

NOTES

1. The heat of formation of Fe_2O_3 is -196.5 kcal/mol.
2. If you want to prepare a bag and use it later to produce heat, squeeze out all the air after everything has been added and tie it securely. When you want to activate the reaction, open the bag, let in some air, and tie the bag again.
3. Vermiculite is a mineral—actually the hydrous silicate of aluminum, magnesium, and iron. When it is heated, it expands. This expanded material is what we use for insulation.
4. Because more iron than oxygen is in the bag, the oxygen is used up first. The reaction can be regenerated by letting in more oxygen.
5. This chemical hand warmer is similar to the hand warmers sold in sporting goods stores.

QUESTIONS FOR STUDENTS

1. Explain how this is an oxidation-reduction reaction. What is oxidized? What is reduced?
2. Explain what function each ingredient serves.
3. How is this reaction similar to "rusting"?

55. The Acid in Water Puzzle

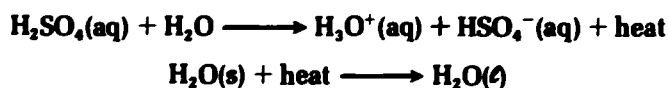
Acid is added to 0 °C water in one cup and 0 °C ice in another cup. The temperature change in each cup is dramatically different, going up in the cup with water and down in the cup with ice.

Procedure

1. Place 100 g of 0 °C water and a thermometer in a Styrofoam-brand cup.
2. Place 100 g of 0 °C ice and a thermometer in another cup.
3. Have two 100-mL portions of room-temperature 9 M sulfuric acid ready.
4. Pour one of the acid portions into each cup.
5. Note the temperature changes in each cup.

Reactions

The heat of solution of acid in water produces the temperature increase in the first cup. The freezing point depression of solution causes the temperature in the second cup to continue to go down to about -15 °C. Any heat generated by the hydration of the acid is absorbed by the phase change of the ice molecules.



Solution

The sulfuric acid concentration is 9 M: Dilute the concentrated acid by placing 100 mL of water in a beaker or flask and adding acid to produce a total volume of 200 mL.

Teaching Tips

NOTES

1. Involve students by having them hold and read the thermometers.
2. Ask the class to predict the result of adding the acid portions to the cups.
3. The acid added to water will increase the temperature 18–20 °C.
4. The acid added to ice will lower the temperature by about 15 °C.
5. The entropy factor is large in the second cup because, in addition to the increased solution, the water molecules are changing from solid to liquid.
6. Be sure to prepare the 9 M acid a few hours in advance so it can cool to room temperature.

QUESTIONS FOR STUDENTS

1. What is *entropy*?
2. Explain how entropy is increasing in each cup.
3. What is the difference in the temperature changes when comparing the two cups?
4. What reason can you give for the difference in the temperature changes?

56. Flaming Cotton

A small amount of a substance is sprinkled over several balls of cotton. When 1 drop of water is added, the cotton balls burst into flame.

Procedure

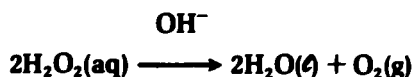
1. Select about six to eight small cotton balls and spread them apart to increase their surface area.
2. Place the cotton on a fireproof or metal surface.
3. Place about 1 tsp of FRESH sodium peroxide on the cotton.
4. CAREFULLY add 1 drop of water from a dropper to the sodium peroxide.
5. Observe the exothermic reaction and production of flame that consumes the cotton.

Reactions

1. Sodium peroxide is a very good source of oxygen gas. When water is added, sodium peroxide reacts to give hydrogen peroxide.



2. The hydroxide ion, OH^- , acts as a catalyst to immediately decompose the hydrogen peroxide to form oxygen gas and water.



3. The heat produced by the reaction, the high oxygen content, and the low ignition temperature of cotton produce the flame.

Teaching Tips

NOTES

1. The sodium peroxide must be fresh. Store the material in a tightly sealed jar.
2. Do not add too much water—only 1 or 2 drops is needed.
3. Most of the sodium peroxide produced in this country is used as a bleaching agent in the textile industry.
4. Sodium peroxide is formed when sodium metal burns in air.

QUESTIONS FOR STUDENTS

1. What is a *catalyst*?
2. What is the catalyst in this reaction? How is it produced?
3. Why does sodium peroxide often lose its "strength" when stored for a long period? (It reacts with water to form sodium hydroxide.)
4. Why does the cotton burn?

57. Nitrocellulose

A small ball of nitrocellulose, which has the appearance of cotton, is touched with a hot glass rod. The ball suddenly flashes and is immediately consumed.

Procedure

1. Place a small ball of prepared nitrocellulose (see Material) on the iron base of a ring stand.
2. Heat a 12-in. glass stirring rod in a burner flame.
3. Carefully touch the hot end of the glass rod to the nitrocellulose—FROM A SAFE DISTANCE!
4. The nitrocellulose will immediately ignite and burn with a flash.

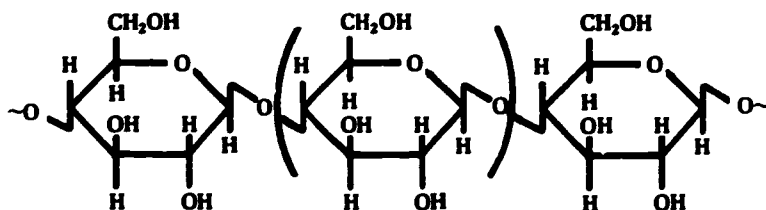
Material

The nitrocellulose is prepared as follows:

1. Place about a dozen cotton balls (which are composed primarily of cellulose) in a small beaker.
2. Carefully combine 50 mL of concentrated nitric acid, HNO_3 , and 100 mL of concentrated sulfuric acid, H_2SO_4 , and pour into the beaker containing the cotton balls. Place the beaker in an ice bath in a hood, cover it with a watch glass, and allow it to sit overnight. **USE EXTREME CARE WHEN HANDLING CONCENTRATED NITRIC AND SULFURIC ACIDS. IMMEDIATELY WASH ANY SPILLS THAT OCCUR.**
3. Rinse the material well with water several times and allow the material to dry.
4. Store the prepared nitrocellulose balls in a closed plastic jar or bag.

Reactions

1. Preparation of the nitrocellulose: Cellulose is a polysaccharide found in plants. Most plant material, like cotton and rayon, are practically all cellulose. Cellulose is similar to starch in its structure.



Concentrated sulfuric acid acts as a dehydrating agent to remove OH groups on the cellulose molecule, and nitric acid nitrates the cellulose and forms the highly reactive nitrate group. Because three nitrate groups are on each unit in the cellulose molecule, it is more properly called *trinitrocellulose*.

2. Combustion of trinitrocellulose: The small amount of heat from the glass rod is sufficient to cause the trinitrocellulose to react and form nitrogen gas as one of the primary products.

Teaching Tips

NOTES

1. Nitroglycerin is a product similar to nitrocellulose. Nitroglycerin is formed by nitrating glycerin.
2. The prepared trinitrocellulose can be safely stored, but not in a glass container. Always keep the plastic jar or bag containing extra material sealed, to prevent sparks or heat from igniting it.
3. Acids remaining after the reaction with cellulose can be flushed down the sink.
4. Nitrocellulose is sometimes called *gun cotton*.

QUESTIONS FOR STUDENTS

1. What is actually "burning" in this reaction?
2. Structurally, how does cellulose differ from starch? (Go to the library!)
3. How would this reaction occur differently if it were ignited in a closed container? (Rapid buildup of nitrogen and oxygen gases would cause an explosion.)

58. The Self-Lighting Candle

A glass stirring rod is touched to the top of a candle. In a few seconds, a flame is produced that lights the candle.

Procedure

1. Make a small depression in the top of a household candle, next to the wick.
2. Fill the depression with the mixture prepared as described under Solution and Material.
3. With the candle firmly secured to its holder, carefully touch the top of the mixture with a long glass stirring rod that has a single drop of sulfuric acid on the tip. BE CAREFUL WITH THE ACID.
4. In a few seconds, the flame will be produced and the candle will appear to light, as if by "magic".

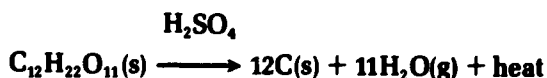
Solution and Material

1. The sulfuric acid is concentrated. USE EXTREME CAUTION WITH CONCENTRATED SULFURIC ACID. YOU ONLY NEED 1 DROP ON THE END OF THE GLASS ROD.
2. The mixture is made by GENTLY mixing a pea-size amount of potassium chlorate with the same amount of table sugar. CAUTION: NEVER GRIND ANYTHING WITH POTASSIUM CHLORATE. IT IS A POWERFUL OXIDIZING AGENT AND MAY EXPLODE. GENTLY MIX THE SUGAR AND POTASSIUM CHLORATE BY PLACING THEM IN THE MIDDLE OF A SHEET OF PAPER AND ROCKING THEM BACK AND FORTH.

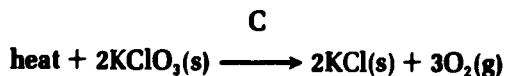
Reactions

This demonstration involves a series of rapid reactions.

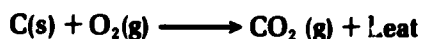
1. Dehydration of the sugar by sulfuric acid produces both heat and carbon, which acts as a catalyst and as a reactant.



2. Heat from reaction 1 initiates the decomposition of potassium chlorate, which proceeds even more rapidly in the presence of carbon, which acts as a catalyst to produce oxygen gas.



3. In the presence of a high concentration of oxygen, the carbon burns rapidly, producing enough heat to vaporize the candle wax and ignite the wick.



Teaching Tips

NOTES

1. Most candles are made of paraffin. A typical reaction for the burning of paraffin is



2. Do not store any unused mixture. Dispose of any left over by flushing it down the drain with copious amounts of water.
3. This demonstration is an effective way to begin a study of thermochemistry.

QUESTIONS FOR STUDENTS

1. Because oxygen is present in the air, why must potassium chlorate be used to generate oxygen? (For the reaction to occur rapidly enough to produce heat to light the candle, a high concentration of oxygen must be present.)
2. What is the source of heat for the decomposition of potassium chlorate?
3. What are three sources of carbon that react rapidly in this series of reactions? (Sugar, carbon from sugar, and wax.)

59. Electrochemical Energy in a Flash

A flash cube has two wires attached by alligator clips. One wire is attached by another clip to a coil of copper wire, and the other wire is attached by a clip to a coil of magnesium ribbon. When the magnesium ribbon and copper wire are dipped into a beaker of acid, the bulb flashes.

Procedure

1. Clean a 12-cm piece of magnesium ribbon by dipping it quickly into a beaker of dilute hydrochloric acid.
2. Clean a 12-cm piece of heavy copper wire by dipping it quickly into a beaker of dilute nitric acid.
3. Using a wire lead with an alligator clip on each end, attach the magnesium ribbon to the wire leads on the metal base of a flash cube with an alligator clip.
4. With another wire lead, attach the copper wire to the contact point at the base of the flash cube.
5. Clamp the flash cube in position on a ring stand.
6. Fill a 250-mL beaker two-thirds full with hydrochloric acid.
7. Raise the beaker into the two metals.
8. The flash cube will flash.

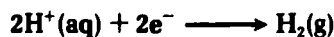
Reactions

This electrochemical cell produces enough voltage to cause the flash cube to flash.

1. At the anode:



2. At the cathode:



The reaction produces, theoretically, 2.37 V. Very little energy is required to set off the bulb, and this is accomplished by the surge of electrons from the reaction.

Solution

The hydrochloric acid concentration is 6 M: See Appendix 2. Sulfuric acid (CAREFUL!) can also be used.

Teaching Tips

NOTES

1. You can also use flash bulbs. Construct a socket for the flash bulb with wire leads soldered on. Attach the copper wire and magnesium ribbon to the leads.
2. Oxidation occurs at the anode; reduction occurs at the cathode.
3. Try to flatten the coiled magnesium and copper so that a maximum surface area is exposed at one time to the acid.
4. Corroded leads, contact points, or electrodes will not produce a flash.

QUESTIONS FOR STUDENTS

1. Explain why the bulb flashed.
2. Is this an oxidation-reduction reaction? What is oxidized? What is reduced?
3. What purpose does the copper serve?
4. Would a weaker acid work? Try it!

60. Chalk That Glows in the Dark

A specially prepared piece of blackboard chalk is used to write a message on a large piece of cardboard in a darkened room. The message glows as it is written.

Procedure

1. Prepare the chalk by soaking several pieces in the solution described under Solution. Allow the chalk to dry.
2. Obtain a large piece of cardboard and cover it with white butcher or wrapping paper.
3. Arrange an ultraviolet (UV) light so that it is directed on the paper.
4. With the lights off and the UV light on, write your message on the paper with the prepared chalk.
5. Enjoy!

Reaction

Anthracene has the formula $C_{14}H_{10}$. It consists essentially of three benzene rings joined together.

As in any reaction that produces light, the UV light excites the anthracene molecule and thus moves electrons to a higher energy level. When the electrons return to their original energy level, energy is released as light.

Solution

Prepare the fluorescent solution by dissolving a pea-size amount of anthracene in 20 mL of toluene or other available organic solvents. Place this solution in a test tube, drop in a piece of chalk, and stopper the tube. CAUTION: USE ONLY SMALL AMOUNTS OF ANTHRACENE AND KEEP THE SOLUTION FOR REUSE.

Teaching Tips

NOTES

1. Other hydrocarbons, such as machine oil and petroleum jelly, will also fluoresce with UV light.
2. As an added precaution, you might want to use a plastic glove when handling the chalk. Anthracene is a suspected mutagen in some animal studies.
3. Be careful to point the UV light away from students and avoid looking directly into it.
4. Magic shops and novelty stores sell safe UV lights, with radiation in the low-UV range.

QUESTIONS FOR STUDENTS

1. What is the difference between ordinary light and UV light? [UV light has more energy and shorter wavelengths than ordinary light. It is thus able to excite electrons that are not affected by ordinary (visible) light.]
2. How is the atom, or molecule, affected by UV light in this demonstration?
3. What are some applications of fluorescence?

61. Chemiluminescence: Glowing School Colors

Solutions are prepared in two large round-bottomed flasks. As two students swirl these flasks in a darkened room, the brightly colored solutions glow for several seconds.

Procedure

1. Select the dyes from the list under Solutions to produce your school colors.
2. Obtain two large (at least 1 L) round-bottomed flasks with stoppers to fit.
3. Add enough dichloromethane to give a volume of about 1 in. to each flask.
4. Add a pinch (pinhead size) of the selected dye to each flask.
5. Add about 0.5 mL of oxalyl chloride to each flask.
6. When ready to produce the chemiluminescence, ask two students to hold the flasks.
7. Simultaneously add a dropper full of hydrogen peroxide to each flask.
8. Immediately stopper each flask, lower the lights, and swirl the flasks to produce the desired colors.

Reactions

1. When hydrogen peroxide reacts with oxalyl chloride, an intermediate product is formed that decomposes to release energy.
2. The energy is transferred to the dye molecule and excites electrons to a higher energy level. As the electrons return to the ground state, visible light is emitted.
3. The color of light emitted (wavelength) depends upon the chemical structure of the dye.

Solutions

1. Hydrogen peroxide: 30% hydrogen peroxide works best. CAUTION! Handle this solution only with plastic gloves. The 3% (drugstore variety) may work, but a larger volume is needed.
2. Dichloromethane, CH_2Cl_2 , is an organic solvent. Handle this material with care, because it is a suspected human carcinogen.
3. Oxalyl chloride should be used in the hood. Wear gloves when dispensing this material.
4. Dyes. Select the dye to give the color desired:

Dye	Color
Perylene	Blue
Tetracene	Green
Rubrene	Yellow
Isoviolanthrone	Orange
Violanthrone	Red

Teaching Tips

NOTES

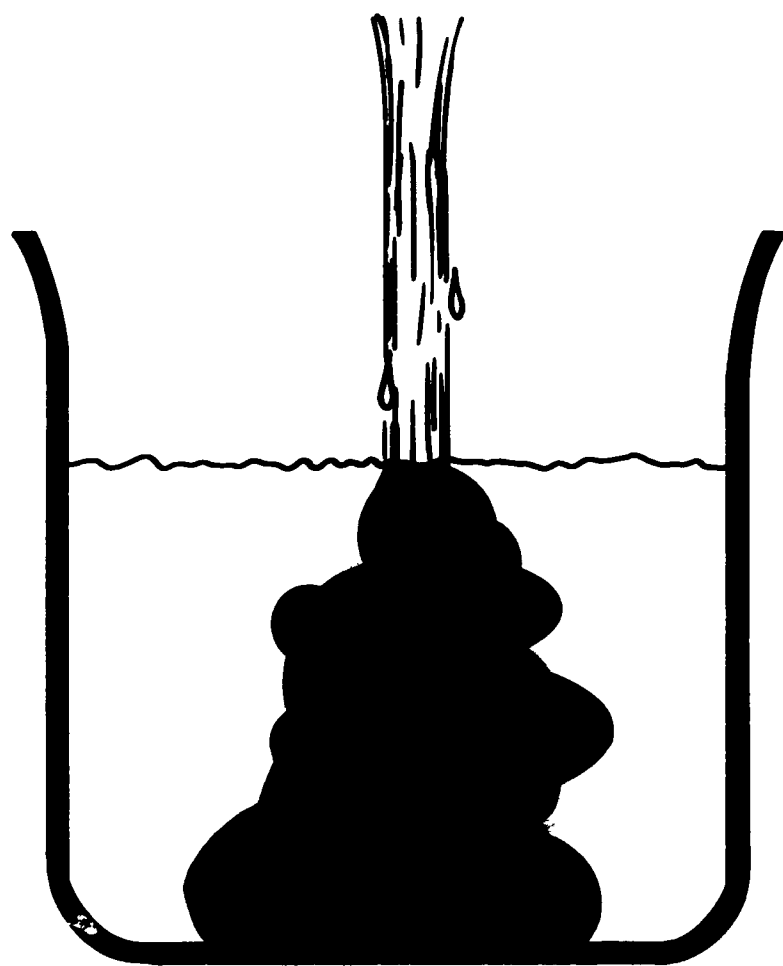
1. These dyes are inexpensive. They can be purchased from any chemical supply company. A good source is Aldrich Chemical Co. Handle them with care, as you would any polycyclic hydrocarbon. They are also suspected carcinogens.

2. Oxalyl chloride, $(-\text{CONH}_2)_2$, can also be ordered from any chemical supply company.
3. These dyes are complex structures and usually consist of a large number of benzene rings.
4. Try varying amounts of reagents and dyes until you obtain the desired colors and duration of chemiluminescence.
5. If smaller containers must be used, adjust the amounts of reagents accordingly.
6. Oxalyl chloride has the structure $\text{Cl}-\text{C}-\text{C}-\text{Cl}$.
 $\begin{array}{cc} \parallel & \parallel \\ \text{O} & \text{O} \end{array}$

QUESTIONS FOR STUDENTS

1. How can you explain the production of light in this demonstration?
2. Why do different dyes produce different colors?
3. What are the structures of some of these dyes? (Go to the library!)
4. What does hydrogen peroxide do in this reaction?

Solutions and Solubility



62. Chemical "Miracles" from 1808

Many reactions fascinated early chemists. These two gel formations were described in a textbook, *The Chemical Catachism*, written in 1808, only 5 years after John Dalton proposed the atomic theory of matter.

Procedure

1. To prepare the first gel place 25 mL of calcium chloride solution in a large test tube and slowly add 25 mL of potassium carbonate solution to the tube. Stir.
2. To prepare the second gel place 25 mL of magnesium sulfate solution in a large test tube and slowly add 25 mL of sodium or potassium hydroxide solution to it. Stir.

Reactions

1. In the first reaction, the calcium ions react with the carbonate ions to form a hydrated product, calcium carbonate.



2. In the second reaction, the magnesium ions react with the hydroxide ions to form a hydrated magnesium hydroxide.



Solutions

1. Calcium chloride, CaCl_2 : The saturated solution is 75 g/100 mL of solution.
2. Potassium carbonate, K_2CO_3 : The saturated solution is 112 g/100 mL of solution.
3. Magnesium sulfate, MgSO_4 : The saturated solution is 26 g/100 mL of solution.
4. Potassium hydroxide, KOH : The saturated solution is 107 g/100 mL of solution.
5. Sodium hydroxide, NaOH : The saturated solution is 42 g/100 mL of solution.

Teaching Tips

NOTES

1. These reactions were first described in *The Chemical Catachism, with Notes, Illustrations, and Experiments* (Parkes, S. 3rd ed.; Lackington, Allen, and Company: London, 1808).
2. On page 561 are the following procedures:

Mix in a wine glass equal quantities of a saturated solution of muriate of lime and a saturated solution of carbonate of potassium, both transparent fluids: stir the mixture, and a SOLID MASS will be the product.

Take a transparent saturated solution of sulphate of magnesia (Epsom salt) and pour into it a like solution of caustic potash, or soda. The mixture will immediately become almost SOLID. The instances of the sudden conversion of two fluids to a solid have been called *chemical miracles*.

3. A gel is a colloid in which the particles link together with some strength. In this case, hydrogen bonding of water molecules, which hydrate the precipitates, probably produces the strong linkages.
4. In each case of the formation of the original solutions, the process is exothermic. You might want to use thermometers and prepare the solutions in the presence of the students. Cool the solutions before using. High energies of hydration during solution formation indicate strong attractions of the ions with water molecules.
5. Try adding an acid to each of the gels. The gels will dissolve, and the observations will provide students with an additional opportunity to identify reaction products and practice writing equations.
6. Point out that students in 1808 were only required to observe and describe the reactions because theory on the particle level was not known, systematic names and symbols for substances had not been devised, and the concept of the ion awaited the discovery of the electron 100 years later!

QUESTIONS FOR STUDENTS

1. Imagine yourself a student in 1808. Write a short description of what you think is happening when the two solutions are mixed in each case.
2. Describe the gels. What is distinctive about them in comparison with ordinary precipitates?
3. Write equations for the dissolving of each of the four original solids.
4. Write equations for the formation of each of the two solid gels.
5. Write equations for the reactions of the gels with an acid.

63. Ions in Slow Motion

Small amounts of two solutions are placed close to opposite sides of a Petri dish with a water pool between them, and the dish is projected on an overhead projector. Merging of the two solutions with the water pool causes slow movement of the ions with the gradual appearance of a haze that darkens.

Procedure

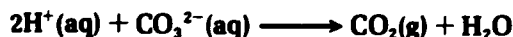
1. Using a medicine dropper, place a small amount of 0.1 M copper sulfate solution close to one side of a Petri dish. Place a small amount of ammonium hydroxide solution close to the other side.
2. Using a dropper, place a small water pool between the two solutions but not touching either one.
3. With a stirring rod, carefully merge the two solutions with the water pool. Avoid mechanical mixing.
4. The two solutions will approach each other slowly and form a gradually darkening haze as the precipitate continues to form.
5. Repeat this demonstration using sodium carbonate solution and 3 M sulfuric acid. This reaction will result in bubbles of carbon dioxide being formed.

Reactions

1. Copper sulfate and ammonium hydroxide react to form copper(II) hydroxide.



2. Sulfuric acid and sodium carbonate react to form carbon dioxide gas.



Solutions

1. The copper sulfate solution concentration is 0.1 M: Dissolve 2.5 g of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in 100 mL of solution.
2. The ammonium hydroxide solution concentration is 0.1 M: See Appendix 2 for dilution instructions.
3. The sodium carbonate solution concentration is 1.0 M: Dissolve 10.6 g of Na_2CO_3 in 100 mL of solution.
4. The sulfuric acid concentration is 3 M: See Appendix 2 for dilution instructions.

Teaching Tips

NOTES

1. This demonstration is best done by using an overhead projector. The darkening haze and carbon dioxide bubbles show up very nicely.
2. Lead nitrate and potassium chromate can be substituted for the copper sulfate and ammonium hydroxide. Other combinations that produce precipitates would also work.
3. You need at least a 1.0 M sodium carbonate solution to get a good production of carbon dioxide bubbles. A higher concentration than 1.0 M would even be better.

4. An interesting sidelight might be to generate enough of the carbon dioxide in a test tube by water displacement to do the limewater (calcium hydroxide) test. This test would help verify that the product is carbon dioxide.

QUESTIONS FOR STUDENTS

1. By what process do the solutions mix? (Diffusion.)
2. What is combining to cause the reaction?
3. What gas is being generated? Test it to verify.
4. Write the equations for the reactions.

64. Supersaturation

A flask containing a liquid is shown to the class. The stopper is removed from the flask, and a single solid crystal is added. Immediately, the entire contents of the flask solidify and the flask becomes warmer.

Procedure

1. Prepare a flask as described under Solution.
2. Present the flask containing a liquid to the class.
3. Carefully remove the stopper from the flask and add *one* tiny crystal of solid sodium acetate trihydrate.
4. Stopper the flask. Observe as the liquid in the top part of the flask begins to solidify and the solidification increases throughout the flask.
5. Note that the flask has become warmer.

Reactions

The supersaturated solution of sodium acetate trihydrate holds more dissolved solute than would normally be in equilibrium with undissolved solute. When this equilibrium is suddenly disturbed by adding a small crystal of solute, by sudden cooling, or even by scratching the container, all of the excess solute precipitates from solution.

Solution

Fill a clean, dry round-bottomed flask with solid sodium acetate trihydrate ($\text{CH}_3\text{COONa} \cdot 3\text{H}_2\text{O}$). Slowly heat the flask on a hot plate until the material completely liquifies.

Heat for a few minutes but do not boil. Remove the flask from the heat and carefully rinse down the neck with a small amount of water from a wash bottle. Insert a stopper and allow the flask to cool at room temperature.

Teaching Tips

NOTES

1. You can use the same flask over and over, just heat after each use to liquify the crystals.
2. You can add small crystals of other substances, such as copper sulfate, to induce solidification.
3. The flask becomes warm because of heat of crystallization of sodium acetate.
4. You can also use sodium thiosulfate.
5. A supersaturated solution of acetanilide, when cooled, will form crystals when the temperature is lowered to 50°C or when a small crystal is introduced.
6. Honey is a good example of a supersaturated solution. Some water evaporates in the beehive as honey is being made. When honey stands on the shelf for a long time, excess sugar may precipitate.

QUESTIONS FOR STUDENTS

1. Describe the reaction. Was heat evolved? Why?
2. What would happen if a crystal of another substance (copper sulfate, for instance) was added instead of sodium thiosulfate?
3. What function does the added crystal serve?

65. Name That Precipitate

Twelve large test tubes are arranged on the demonstration desk. As the solutions are mixed, two at a time, various precipitates are formed. These brown, red, white, blue, and yellow precipitates vividly show various ionic reactions.

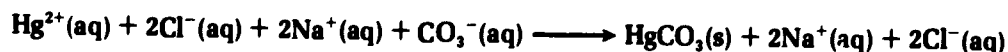
Procedure

1. Arrange 12 test tubes in a test-tube holder.
2. Label the tubes as follows:

1, HgCl ₂	7, HgCl ₂
2, Na ₂ CO ₃	8, KI
3, NiCl ₂	9, BaCl ₂
4, Na ₂ CO ₃	10, Na ₂ SO ₄
5, Pb(NO ₃) ₂	11, CuCl ₂
6, KI	12, Na ₂ CO ₃
3. Fill each test tube one-third full of its solution.
4. Pour the contents of test tube 1 into test tube 2. What do you observe?
5. Ask students to write an *ionic* equation for this reaction.
6. Ask students to name the precipitate.
7. Mix the test tubes as follows. In each case, have students describe the reaction and predict the formula for the precipitate.
 - a. Test tube 3 into test tube 4.
 - b. Test tube 5 into test tube 6.
 - c. Test tube 7 into test tube 8.
 - d. Test tube 9 into test tube 10.
 - e. Test tube 11 into test tube 12.

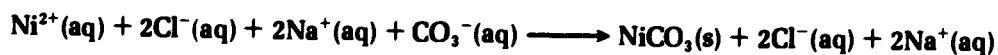
Reactions

1. Mercury(II) chloride and sodium carbonate produce the BROWN precipitate, mercury carbonate.

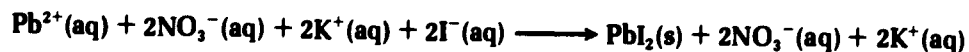


(Ions, such as Cl and Na, that do not enter a reaction are often called *spectator* ions.)

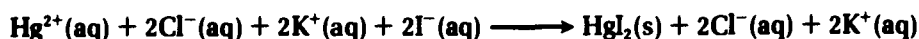
2. Nickel chloride and sodium carbonate produce the pale BLUE precipitate, nickel carbonate.



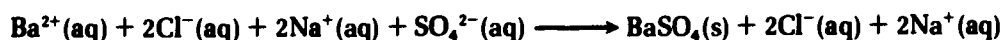
3. Lead nitrate and potassium iodide produce the YELLOW precipitate, lead iodide.



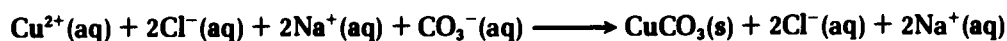
4. Mercury(II) chloride and potassium iodide produce the ORANGE-RED precipitate, mercury iodide.



5. Barium chloride and sodium sulfate produce the WHITE precipitate, barium sulfate.



6. Copper chloride and sodium carbonate produce the pale BLUE precipitate, copper carbonate.



Solutions

Make dilute solutions of each of the following by adding about 1 tsp of each to 500 mL of water.

1. Mercury(II) chloride, HgCl_2 .
2. Sodium carbonate, Na_2CO_3 .
3. Nickel chloride, NiCl_2 .
4. Sodium carbonate, Na_2CO_3 .
5. Lead nitrate, $\text{Pb}(\text{NO}_3)_2$.
6. Potassium iodide, KI .
7. Barium chloride, BaCl_2 .
8. Sodium sulfate, Na_2SO_4 .
9. Copper(II) chloride, CuCl_2 .

Notice that some of the solutions [potassium iodide, sodium carbonate, and mercury(II) chloride] are used in more than one combination of solutions.

Teaching Tips

NOTES

1. Precipitates with characteristic colors and appearances result when certain ionic solutions are mixed. Recombination of certain positive and negative ions often results in the formation of a compound with low solubility. That compound will precipitate as a solid.
2. A precipitate will form in each case. The students will then "deduce" the formula of the precipitate. You should have solutions of NaCl , KCl , and KNO_3 available for them to see, as evidence that these compounds are soluble in water and would not appear as a precipitate.
3. You might try putting these solutions in dropper bottles and letting students mix drops of each on a glass plate.
4. Draw a large chart on the blackboard and have students enter data (colors, name, and formula of precipitate) for each reaction.
5. Positive ions are *cations*, and negative ions are *anions*. Notice that ionic reactions show only ions in solution.
6. See Appendix 4 for disposal of mercury, lead, and barium salts.
7. You might prefer, as the authors did, to use complete equations here, rather than net ionic equations.

QUESTIONS FOR STUDENTS

1. What is an *ionic* solution?
2. What causes a precipitate to form when ionic solutions are mixed?
3. What does this property have to do with solubility?
4. If copper chloride, CuCl_2 , and sodium hydroxide are mixed, a blue precipitate is formed. What is this precipitate? Show the ionic reaction. $[\text{Cu}^{2+} + 2\text{Cl}^- + 2\text{Na}^+ + 2\text{OH}^- \longrightarrow \text{Cu}(\text{OH})_2(\text{s}) + 2\text{Cl}^- + 2\text{Na}^+]$. The precipitate must be $\text{Cu}(\text{OH})_2$, because the only other possible new combination of ions would produce NaCl , table salt, which is soluble in water.]

66. Silver Ion Solubilities: Red and White Precipitates

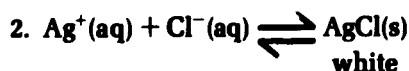
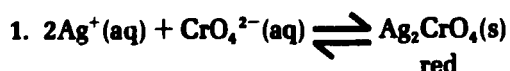
A few drops of a solution are added to a second solution, and a deep red precipitate forms. When a few drops of a third solution are added to the same solution, a white precipitate forms. As more solution is added, the red precipitate dissolves and leaves only the white precipitate.

Procedure

1. Place 200 mL of silver nitrate solution in a 250-mL Erlenmeyer flask or a large beaker.
2. Add 2 drops of the yellow sodium chromate solution. Note the immediate formation of a RED precipitate.
3. Add, dropwise, 1-2 mL of sodium chloride solution. Notice the formation of a WHITE precipitate.
4. After observing and discussing the two precipitates, add 50 mL of sodium chloride solution. Observe that the red precipitate dissolves and leaves only the white precipitate.

Reactions

This demonstration involves the formation of two precipitates from silver nitrate solution.



Because the solubility of red silver chromate is greater than that of silver chloride, more white silver chloride is formed as more chloride is added.

Solutions

1. The silver nitrate concentration is 0.05 M: Dissolve 0.85 g of AgNO_3 in 100 mL of solution.
2. The sodium chloride concentration is 0.2 M: Dissolve 1.2 g of NaCl in 100 mL of solution.
3. The sodium chromate concentration is 0.5 M: Dissolve 8.1 g of Na_2CrO_4 in 100 mL of solution.

Teaching Tips

NOTES

1. The solubility of silver chromate is 1.4×10^{-3} g/100 mL of water. The solubility of silver chloride is 8.9×10^{-4} g/100 mL of water.
2. This demonstration offers an effective way to introduce equilibrium. As more chloride is added, more silver ions react. This reaction shifts the equilibrium from the solid silver chromate to form more silver ions and chromate ions. This shift should produce a slight yellow (chromate) tint to the solution.

3. See Appendix 4 for disposal procedures.
4. Allow the precipitate to settle for an hour or so to make the extent of precipitation more obvious.

QUESTIONS FOR STUDENTS

1. Explain how each precipitate forms in this demonstration.
2. Why does the red precipitate dissolve?
3. Is there a slight yellow color in the final solution? If so, why?
4. Theoretically (but not practically), how would you dissolve the white precipitate?

67. Patriotic Precipitates

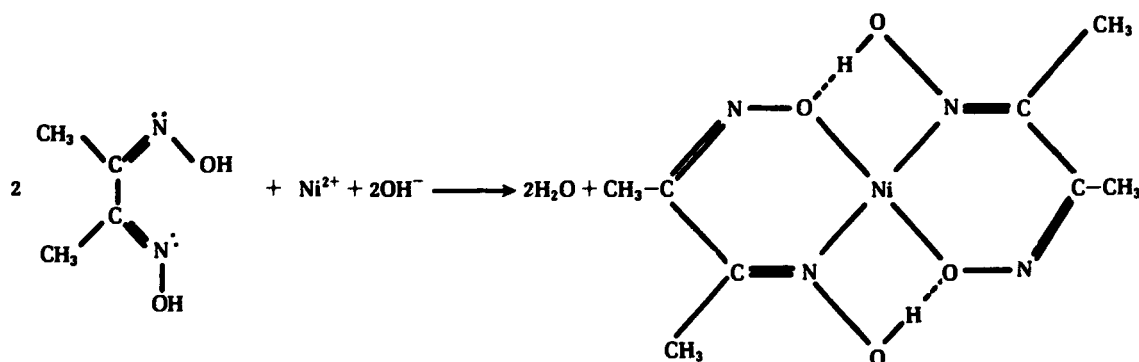
Solutions are mixed to immediately form RED, WHITE, and BLUE precipitates.

Procedure

1. **RED precipitate:** Add equal amounts of nickel sulfate solution and dimethylglyoxime solution. Add a few drops of concentrated ammonium hydroxide if a deep red color does not immediately form.
2. **WHITE precipitate:** Add equal amounts of calcium chloride solution and potassium carbonate solution.
3. **BLUE precipitate:** Add equal amounts of iron(II) sulfate solution and potassium ferricyanide solution.

Reactions

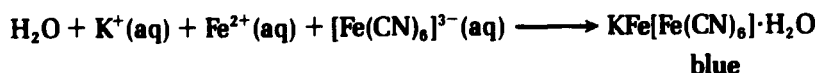
1. Nickel ion reacts with dimethylglyoxime to form the red precipitate, nickel dimethylglyoxime.



2. Potassium carbonate and calcium chloride react to form the white precipitate, calcium carbonate.



3. Iron sulfate reacts with potassium ferricyanide to produce a blue complex.



Solutions

1. **RED precipitate:** Dissolve 5 g of nickel sulfate (NiSO_4) in 100 mL of water and 5 mL of dimethylglyoxime in 100 mL of ethyl alcohol.
2. **WHITE precipitate:** Dissolve 5 g of calcium chloride (CaCl_2) in 100 mL of water and 10 g of potassium carbonate (K_2CO_3) in 100 mL of water.
2. **BLUE precipitate:** Dissolve 2 g of iron(II) sulfate (FeSO_4) in 100 mL of water and 3 g of potassium ferricyanide in 100 mL of water.

Teaching Tips

NOTES

1. Try varying the concentrations of the solutions until the desired intensity of color results.
2. This demonstration is a good review of chemical reactions!
3. Dimethylglyoxime is an organic chelating agent. A *chelate* is a ligand with two or more donor atoms. When dimethylglyoxime reacts with nickel, it forms a metal chelate. Hemoglobin in blood contains the complex heme, a square-planar complex of iron. Chlorophyll has a square-planar magnesium complex.
4. The scarlet complex that forms with dimethylglyoxime and nickel ion is a characteristic test for nickel. Only palladium gives a similar color.
5. Reactions 1 and 3 are used to identify the ions Ni^{2+} and Fe^{2+} in standard qualitative analytical procedures.
6. Calcium carbonate is a form of chalk.
7. Potassium ferricyanide is more properly called potassium hexacyanoferrate(III). The blue product, $\text{K}_3\text{Fe}(\text{CN})_6 \cdot \text{H}_2\text{O}$ results when iron(II) reacts with hexacyanoferrate(III), or when iron(III) reacts with hexacyanoferrate(II).

QUESTIONS FOR STUDENTS

1. Think of another colored precipitate and a reaction to form it.
2. Write the net ionic equation for the formation of CaCO_3 .
3. What colors do you expect the solutions to be if the precipitates are allowed to settle overnight? Try it!

68. A Glittering Shower of Lead Iodide Crystals

A flask containing a warm solution of lead iodide is prepared. As the solution cools, beautiful crystals of lead iodide form and fall to the bottom in a golden shower.

Procedure

1. Fill a large flask two-thirds full with water and heat to boiling.
2. Add solid lead iodide, small amounts at a time, until no more will dissolve.
3. Stopper the flask and allow it to cool.
4. Observe that on cooling yellow crystals of lead iodide form throughout the flask and eventually fall to the bottom in a golden shower.

Reaction

Lead iodide has a low solubility in water at room temperature but is more soluble in hot water.

Teaching Tips

NOTES

1. The solubility of lead iodide is 0.41 g/100 mL in hot water and 0.044 g/100 mL in cold water.
2. The cooled flask can be stoppered and reused many times by reheating the solution to boiling to redissolve the crystals.
3. You can, of course, speed up the crystallization by cooling the flask.
4. Lead iodide is easily prepared by combining saturated solutions of potassium iodide and lead nitrate. Filter the precipitate and add it to the hot water as in step 2 under Procedure.

QUESTIONS FOR STUDENTS

1. What effect does temperature generally have on the solubility of a solid in water?
2. Describe the unique characteristics of the lead iodide crystals.
3. Do the crystals form differently when the solution is cooled slowly as compared with when it is cooled rapidly?

69. Red and White Precipitates in Sodium Silicate

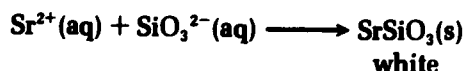
Drops of a clear solution are added to another clear solution, and a white, doughnut-shaped, gelatinous mass is produced. Drops of another clear solution are added, and red doughnut-shaped masses are produced.

Procedure

1. Place about 100 mL of sodium silicate solution in a small beaker.
2. Using a dropper, add solution A dropwise.
3. Note the appearance of a white, gelatinous precipitate.
4. Using another dropper, add solution B dropwise.
5. Note the formation of a red precipitate.

Reactions

1. All of the silicates, except those of Na^+ , K^+ , Rb^+ , Cs^+ , and NH_4^+ , are practically insoluble in water.



2. All soluble silicates form basic solutions. Thus, the red color is due to the reaction of phenolphthalein in a base.

Solutions

1. Sodium silicate solution: Dilute sodium silicate, Na_2SiO_3 , to 2 parts of sodium silicate solution to 5 parts of water.
2. Solution A: strontium chloride solution. Add about 1 tsp of strontium chloride, SrCl_2 , to 50 mL of water.
3. Solution B: strontium chloride. Make as described for solution B but add 3-5 drops of phenolphthalein solution.

Teaching Tips

NOTES

1. Silicates are generally mixtures of species including SiO_3 , SiO_4^{2-} , and SiO_4^{4-} .
2. A water solution of such a mixture of silicates is sold in hardware stores as "water glass".
3. This demonstration projects nicely. Add the drops to sodium silicate in a Petri dish on an overhead projector.

QUESTIONS FOR STUDENTS

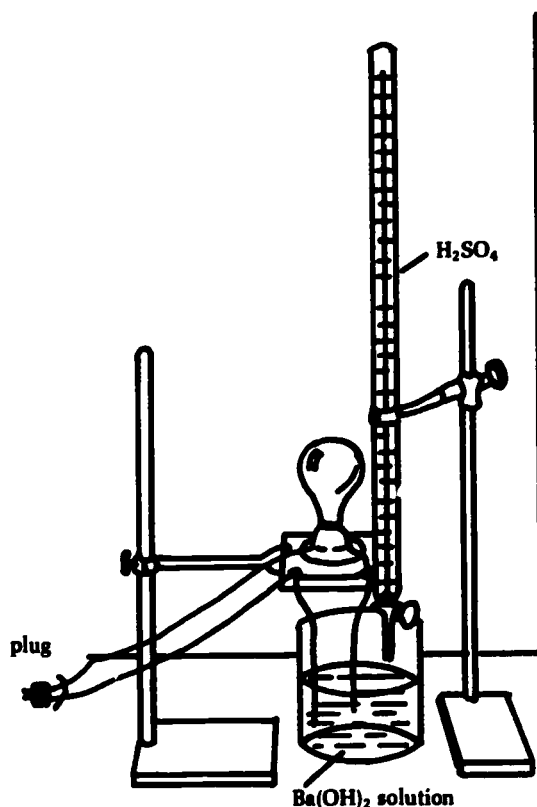
1. What is a *silicate*?
2. What are the precipitates in the demonstration?
3. Why is one precipitate white, and the other is red?

70. Electrolytic Titration

Electrodes from a conductivity apparatus are placed in a pink solution. Acid from a buret is dropped into the solution. The bulb of the conductivity apparatus grows dimmer, the solution becomes cloudy, and the pink color fades. As more acid is added, the bulb suddenly becomes brighter.

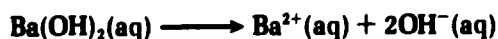
Procedure

1. Set up a conductivity apparatus as shown in the diagram.
2. Place about 200 mL of barium hydroxide solution in a large beaker and place the two electrodes in the beaker. Add 3-5 drops of phenolphthalein solution to the beaker.
3. Position a buret so that the tip is directly over the beaker.
4. Fill the buret with sulfuric acid.
5. Plug in the conductivity apparatus and notice that the bulb glows brightly.
6. Slowly adjust the stopcock on the buret to produce a steady dripping of acid from the buret.
7. Stir the solution in the beaker with a glass stirring rod.
8. Observe.



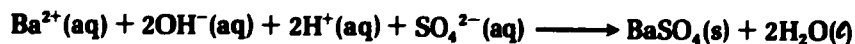
Reactions

1. The light glows when the conductivity apparatus is plugged in because the beaker contains ions in solution.



2. As acid is added, the following is observed:

- a. The light becomes dimmer. This result occurs because fewer ions are available to conduct a current when sulfate ions from the acid react with barium to form the precipitate barium sulfate.



- b. The cloudy appearance is from the barium sulfate precipitate.
- c. The pink color fades because the added hydrogen ion from the acid reacts with the OH^{-} ions in solution and thus lowers their concentration. This decrease in hydroxide ion concentration causes the indicator to lose its color. The H^{+} concentration increases also.
- d. As MORE acid is added, the increased concentration of hydrogen ions and sulfate ions will conduct the current, which causes the bulb to again grow bright.

Solutions

1. Barium hydroxide is a saturated solution. Dissolve 4 g per 200 mL of solution
2. The sulfuric acid concentration is 1.0 M. See Appendix 2.

Teaching Tips

NOTES

1. This demonstration is an excellent way to correlate a great deal of chemistry: neutralization, precipitate formation, acid-base reactions, ionic reactions, etc.
2. Try various concentrations of sulfuric acid to vary the time of the reaction.
3. Be sure to make the solutions with distilled water.
4. In this reaction, the phenolphthalein becomes colorless at pH 8-10, NOT pH 7.
5. If you have a milliammeter, you can use it to show when neutralization occurs.
6. Use a magnetic stirrer if you have one.

QUESTIONS FOR STUDENTS

1. Why is the $\text{Ba}(\text{OH})_2$ solution pink?
2. What causes the bulb to glow brightly in the original solution?
3. What causes the dimming of the bulb? Write the equation for the reaction.
4. Why does the bulb again glow after the pink color is gone?

71. Do Frozen Solutions Conduct Electricity?

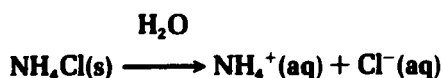
Passing electricity through an aqueous ionic solution causes a bulb to glow. When the same solution is frozen, electricity is not conducted and the bulb does not glow.

Procedure

1. If you do not have a conductivity apparatus, construct one following the procedure given under Solution and Material.
2. Half fill a U-tube with ammonium chloride solution.
3. Place one electrode in each open end of the tube.
4. Notice that the bulb glows; this glow indicates that the ions in solution conduct an electric current.
5. Place the U-tube in a large beaker and fill the beaker with crushed dry ice.
6. Repeat the conductivity test.
7. Notice that the bulb will eventually cease to glow as the temperature drops and the solution freezes.

Reactions

1. The ammonium and chloride ions in solution will conduct a current, and the bulb will glow.



2. When the solution is frozen by the dry ice, the ions are fixed in position and unable to conduct a current; thus, the bulb will cease to glow.

Solution and Material

1. Ammonium chloride: Dissolve 10 g in 100 mL of solution.
2. Make a conductivity apparatus: Connect three 6-V dry cells in series with a flashlight bulb in the circuit. Remove the insulation from the ends of two wires (copper works well) to serve as positive and negative electrodes in the solution.

Teaching Tips

NOTES

1. Try other salt solutions in various concentrations.
2. An even better ice bath can be prepared by making a slurry of dry ice and acetone.
3. Dry ice (solid carbon dioxide) has a temperature of -78°C .
4. The smaller the U-tube, the better. The quick drop in temperature may crack larger tubes due to sudden expansion of the freezing solution.
5. If you do not have a U-tube, bend a piece of glass tubing or connect two pieces of glass tubing with a short piece of Tygon tubing.

QUESTIONS FOR STUDENTS

1. Which ion carries a negative charge through the solution?
2. Would you expect a solid crystal of ammonium chloride to conduct an electric current? Try it!
3. Would you expect molten sodium chloride (melting point is about 800 °C) to conduct an electric current?

72. Osmosis and the Egg Membrane

A raw egg is placed in a beaker containing vinegar. After several days, the shell is gone and the semipermeable membrane of the egg is exposed. Osmosis causes the egg to swell. When the enlarged egg is placed in a syrup solution, the reverse process occurs, which results in a much smaller than normal egg.

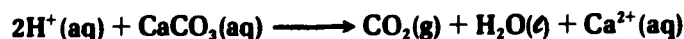
Procedure

1. Place a raw egg in a 250-mL beaker.
2. Fill the beaker with vinegar, and cover it with plastic wrap. Punch a few holes in the wrap.
3. After about 48 h, notice that the egg shell has "disappeared" and that the egg has swollen considerably.
4. Carefully pour off the solution and examine the egg. The membrane is firm enough to allow the egg to be held in the hand.
5. Place the enlarged egg in a second beaker containing a diluted clear syrup solution (50:50).
6. After another 2 days, observe that the egg has "shrunk" to a size much smaller than that of the original egg.

Reactions

This demonstration shows osmosis, the movement of water across a semipermeable membrane from an area of lower to higher solute concentration.

1. When placed in vinegar, the acetic acid reacts with the calcium carbonate in the egg shell.



A very high concentration of protein (mostly albumin) exists in the egg. Thus, water enters the egg in an attempt to make the solute concentration equal on both sides of the semipermeable membrane (egg membrane). Thus, the egg gets larger.

2. When the egg is placed in a syrup solution, a higher solute concentration (glucose) is outside the egg than inside. Thus, water leaves the egg in an attempt to dilute the more concentrated glucose solution outside. Thus, the egg gets much smaller.

Solutions

1. Vinegar is 5% acetic acid.
2. Syrup solution: For best results, use a 50:50 dilution.

Teaching Tips

NOTES

1. Osmosis is responsible for the movement of small nutrient molecules and water into cells, for the movement of waste products (such as urea) out of the cell, for the production of pickles from cucumbers, and so forth.

2. Although this demonstration requires several days, you can set it up ahead of time. When the egg is first placed in vinegar, students will be able to see an immediate reaction with the production of gas bubbles.
3. The pressure that counteracts the movement of water across such a semipermeable membrane is called *osmotic pressure*.
4. This demonstration is a good activity for students to do at home.

QUESTIONS FOR STUDENTS

1. What is *osmosis*? What is *osmotic pressure*?
2. Why does the egg become larger when placed in vinegar?
3. Give the reaction for the dissolving of the egg shell.
4. Why does the egg become smaller when placed in the syrup solution?

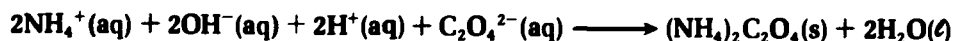
73. Growing Ammonium Oxalate Crystals

A hot solution is filtered, and beautiful white crystals are immediately formed in the filtrate.

Procedure

1. Using a hot plate, dissolve 3.0 g of oxalic acid in 30 mL of hot water at a temperature of about 75 °C.
2. Neutralize the solution with 16 mL of concentrated aqueous ammonia, adding about 0.5 mL at a time. Use blue litmus paper to be sure that the solution is basic.
3. Filter the solution while it is hot.
4. Notice the formation of white crystals in the filtered solution.

Reaction



Solution

Use concentrated ammonium hydroxide (aqueous ammonia).

Teaching Tips

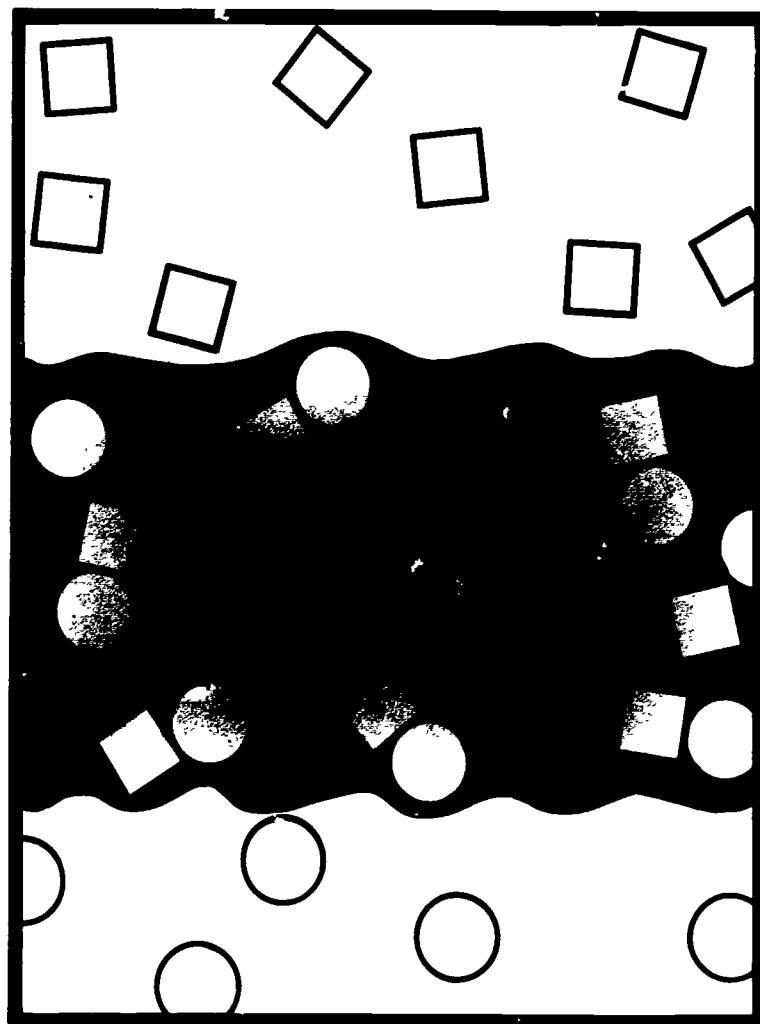
NOTES

1. Use approximately 3 g of oxalic acid for the best crystal formation.
2. Oxalic acid is also called ethanedioic acid.
3. The amount of ammonia added in the procedure ensures an excess of ammonia.
4. Ammonium oxalate is insoluble in an alkaline solution.
5. Adding the ammonia slowly increases the amount of ammonium oxalate formed.

QUESTIONS FOR STUDENTS

1. Why was the solution heated?
2. Explain the reaction that formed the crystals.
3. Why was an excess of ammonia added?
4. Identify the crystals.

Kinetics and Equilibrium



74. Temperature and Reduction of Permanganate

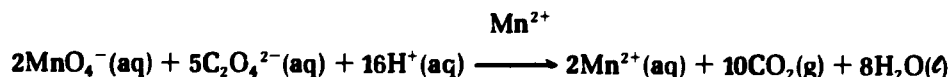
Solutions of purple potassium permanganate and colorless oxalic acid are combined at successively higher temperatures. With each 10-degree increase in temperature, the time for the appearance of a yellow product is cut in half.

Procedure

1. Prepare 12 large test tubes in the following manner: Add 25 mL of 0.005 M KMnO_4 to six test tubes. Add 25 mL of saturated oxalic acid (about 0.75 M) to six test tubes.
2. Determine the room temperature by examining the thermometer. The value is probably around 20 °C.
3. Set up a water bath in a 600–1000-mL beaker. The beaker need be only one-fourth to one-third full of water. Place a thermometer in the water.
4. Place five test tubes of each solution in the water bath. Begin heating to 10 degrees ABOVE room temperature.
5. Meanwhile, watching the time, pour the room-temperature tube of oxalic acid into the room-temperature test tube of potassium permanganate solution. Pour the solution back and forth twice to mix the solutions.
6. Note the exact time that the oxalic acid contacts the potassium permanganate solution. Always use this as the initial time.
7. When the test tube has become yellow, with no hint of pink remaining, note the time. This value is the final time.
8. Place the test tube of reacted solutions in a test-tube rack so that you may observe any additional color change.
9. Repeat this procedure with the reactants at each of the temperatures, which are 10 degrees above the former trial. By watching the water bath closely, you will find this procedure to be quick and efficient.

Reaction

The reaction is the reduction of manganese, which is purple in MnO_4^- (a +7 oxidation state) and very pale pink when the reaction is complete (+2). An intermediate, a yellow intermediate, however, serves as the timed end point for this demonstration because it is more convenient.



Solutions

1. The potassium permanganate concentration is 0.005 M: Dissolve 0.8 g of KMnO_4 in 1 L of solution.
2. Oxalic acid, $\text{H}_2\text{C}_2\text{O}_4$, is saturated (about 0.75 M). Saturated oxalic acid requires 60–70 g/L of solution.

Teaching Tips

NOTES

1. This demonstration should actively involve students: watching the temperature in the bath, pouring and mixing, timing and graphing, and so forth.

2. A graph can be plotted on the board or overhead projector as each new elapsed time value is found. Plot temperature on the x axis and elapsed time on the y axis.
3. Students will quickly see that for each 10-degree increase in temperature the elapsed time for this reaction is cut approximately in half.
4. Typical data are as follows:

Temperature ($^{\circ}\text{C}$)	Time (s)
25	270
35	130
45	60
55	30
65	15
75	7

5. As the Mn^{2+} is produced, in each case it acts as a catalyst, which contributes to the increase in rate; thus, this demonstration is a complex kinetics study.
6. In this demonstration, the students should be told that the temperatures that are recorded are INITIAL temperatures of the reactants because the test tubes are not maintained in a constant-temperature water bath as the reaction proceeds.
7. Expect the room temperature reaction to take 3–4 min. The sixth trial (about 75°C) will, then, take about 7 s.
8. Students can calculate initial rates for each trial by finding $\Delta[\text{MnO}_4^-]/\Delta\text{time}$. The permanganate ion is the limiting reagent because the oxalate is in excess. When the reaction end point is reached, we assume all the permanganate has been consumed. The ΔMnO_4^- concentration, then, is the same as the initial concentration of MnO_4^- . To calculate the initial concentration of permanganate, determine the number of moles of the ion in the 25 mL of 0.005 M solution. Then calculate the molarity in the mixture (2.5×10^{-3} M) by using that number of moles per 50 mL. Rate can then be plotted against temperature to produce a linear plot.

QUESTIONS FOR STUDENTS

1. Compare each reaction time and state a regularity about the relationship between temperature increase and reaction rate increase. (With this reaction, each 10-degree increase in temperature causes a doubling of the rate.)
2. Explain on the particle level why a temperature increase should increase the rate of a reaction.
3. What is the evidence that the permanganate is completely reacted in each trial?

75. Balloon Kinetics

A solid and hydrochloric acid at various temperatures are added to four flasks. Balloons are attached to the mouths of the flasks, and the rate at which the various balloons inflate is a qualitative measure of the reaction rate.

Procedure

1. Inflate four balloons several times to stretch them.
2. Obtain four 250-mL flasks and be sure that you can fit the balloon over the mouth of the flask.
3. Measure four 10-g samples of sodium carbonate.
4. Place a sample in each balloon. Be sure not to lose any of the solid.
5. Prepare four samples of 3 M HCl. Cool one to 0 °C in a ice bath, cool one to 20 °C in cold water, leave one at room temperature, and heat the fourth to 60–70 °C.
6. With four student assistants, carefully cover the mouth of each flask with a balloon. Be careful not to let any of the solid in the balloon enter the flask.
7. When all flasks and balloons have been assembled and all temperatures recorded, simultaneously shake the sodium carbonate from each balloon into the flasks.
8. Observe the rate at which the balloons inflate.
9. Considering the data and observations, draw a conclusion about temperature and rates of reactions.

Reaction

The sodium carbonate and hydrochloric acid produce carbon dioxide gas, which inflates the balloons.



Solution

The hydrochloric acid concentration is 3 M: See Appendix 2.

Teaching Tips

NOTES

1. Several variations are applicable to this demonstration: use vinegar instead of hydrochloric acid and use zinc instead of sodium carbonate to produce hydrogen gas instead of carbon dioxide.
2. Repeat the demonstration, using different concentrations of acid solutions (0.1 M, 1.0 M, 6 M, and concentrated) at the same temperature to see a qualitative effect of concentration on reaction rates.
3. Be sure that the balloons are securely fastened to the flasks.
4. Discuss variables that may affect this demonstration (elasticity of different balloons, etc.).

QUESTIONS FOR STUDENTS

1. Show the reaction for producing the gases in this demonstration.
2. According to the order in which the balloons filled, what generalizations can you make about temperature and reaction rates?
3. What variables would influence the outcome of this demonstration?

76. Appearing Red

A colorless solution is shown to the class and placed on a magnetic stirrer. After a few seconds, the solution turns pink-red in color.

Procedure

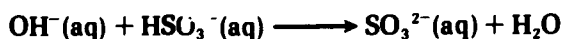
1. Place about 250 mL of water in a large beaker.
2. While stirring, preferably with a magnetic stirrer, add 3-5 drops of phenolphthalein indicator solution.
3. Add a small amount (less than a match-head size) of sodium bisulfite.
4. Add just a few crystals (6-10) of sodium sulfite.
5. When you are ready for the demonstration, add 4-5 drops of formalin to the center of the vortex of the stirring solution.
6. Within a few seconds, notice that the solution will flash a pink-red color. (You may also see a red "tornado" just before the solution changes color.)

Reactions

1. The slow first step of this reaction involves the production of hydroxide by the reaction of formalin (formaldehyde solution) with sulfite.



2. As soon as the OH^- is formed in the first reaction, it is consumed in the second reaction.



Sulfite produced in the second reaction produces more OH^- in the first reaction. When the bisulfite is consumed, hydroxide produced in the first reaction reacts with the phenolphthalein indicator to produce the red color.

Solution

Formalin is a 37% solution of formaldehyde in water.

Teaching Tips

NOTES

1. This reaction can be used effectively to introduce chemical kinetics. Vary the temperature and concentration of reactants and plot them against time.
2. Formalin (not formaldehyde) is used to disinfect and to preserve biological specimens. Its major use is in the manufacture of plastics and adhesives.
3. Formalin should be used with caution. Formaldehyde is a possible carcinogen in animals.
4. Adjust the concentration of reactants to give a deep red color and a time of reaction at room temperature of about 15 s.

QUESTIONS FOR STUDENTS

1. What are the three reactions involved in this demonstration?
2. Which is the slow, or rate-determining, reaction?
3. What causes the red color to appear?
4. Discuss the effects on this reaction of varying the temperature and the concentration of reactants.

77. Disappearing Red

A red solution is stirred on a magnetic stirrer. After a few seconds, the color gradually fades and then completely disappears.

Procedure

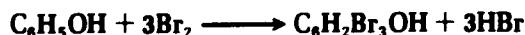
1. Add 40 mL of phenol solution to a large beaker on a magnetic stirrer.
2. Add 80 mL of sulfuric acid.
3. Add 60 mL of sodium bromate solution.
4. Add 6–8 drops of methyl red indicator solution.
5. When you are ready to make the red color disappear, add 30 mL of sodium bromide solution.

Reactions

1. The mechanism for this reaction is a three-step process—one slow step and two fast steps. In the slow first step, bromine is formed.



2. In the second step, aqueous phenol is halogenated to form 2,4,6-tribromophenol.



3. When all of the phenol has reacted, in the final step, aqueous bromine from the first step reacts with the methyl red indicator, to cause the red color to disappear.

Solutions

1. The phenol solution is 0.04%: Dissolve 0.4 g of phenol crystals in 1 L of solution.
2. The sulfuric acid concentration is 0.2 M.
3. The sodium bromate concentration is 0.2 M: Dissolve 30.2 g of NaBrO_3 in 1 L of solution.
4. The sodium bromide concentration is 0.2 M: Dissolve 20.6 g of NaBr in 1 L of solution.

Teaching Tips

NOTES

1. If you do not have phenol crystals, you can use Campeno-Phenique, sold in the drug stores and used to treat fever blisters. Use about a match-head amount from the tube.
2. If the reaction occurs too quickly, dilute the sodium bromide solution.
3. This reaction offers an interesting way to begin the study of chemical kinetics.
4. Phenol was originally called *carbolic acid*. It is used as a disinfectant and in the manufacture of dyes and plastics.

QUESTIONS FOR STUDENTS

1. Why does the red color disappear in this demonstration?

2. Which of the reactions are fast and which are slow? Which reaction determines the rate of the reaction?
3. What is phenol used for?

78. A Variation of the Starch-Iodine Clock Reaction

A colorless solution (100 mL) is added to an equal volume of a second colorless solution in a large beaker. In a few seconds, the solution suddenly turns dark blue. When the temperature is increased, the reaction occurs more rapidly. When one solution is diluted, more time is required for the color to change.

Procedure

1. Place 100 mL of solution A in a beaker.
2. Place 100 mL of solution B in a second beaker.
3. Mix the two solutions by pouring them back and forth twice.
4. Notice the time required for the sudden appearance of a deep blue color.
5. Repeat, using solutions that have been warmed in a water bath to 40-50 °C. Notice that the reaction occurs in a shorter time.
6. Repeat, using solutions that have been cooled in an ice bath. Notice that the time for the reaction is increased.
7. Repeat, using a dilute solution of solution B. Notice the increased time for reaction.

Reactions

This demonstration involves a series of reactions that may be summarized as follows:

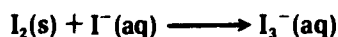
1. Reactions of hydrogen peroxide with iodide ion.



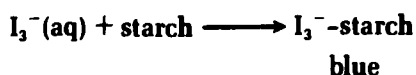
2. As I_2 is formed, it quickly reacts with thiosulfate and re-forms iodide ion.



3. When the thiosulfate is consumed, the I_2 concentration increases and, with the iodide ion, forms the triiodide complex.



4. The triiodide complex ion reacts with starch to form the characteristic blue-black complex.



Solutions

1. Solution A: Prepare 1 L of solution as follows:

- a. Add and stir 0.2 g of soluble starch in 100 mL of boiling water. Heat, but do not boil, for 3-4 min. Dilute to 800 mL.
- b. Add 30 mL of concentrated acetic acid, CH_3COOH .
- c. Add 4.1 g of sodium acetate, NaO_2CCH_3 .
- d. Add 50.0 g of potassium iodide, KI.
- e. Add 4.7 g of sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$.
- f. Allow the solution to cool and dilute to 1.0 L.

2. Solution B: 3% hydrogen peroxide.

Teaching Tips

NOTES

1. This demonstration is a variation of the familiar starch-iodine clock reaction. The advantage is a simpler reaction mechanism.
2. The acetic acid-sodium acetate is added as a buffer to keep the hydrogen ion concentration constant.
3. A small amount of thiosulfate is used so that very little of the H_2O_2 is used. Thus, the concentrations of H_2O_2 , I^- , and H^+ are kept nearly constant.
4. Repeat the demonstration using several temperatures and several dilutions to clearly establish the relationship between temperature and reaction rate and concentration and reaction rate.

QUESTIONS FOR STUDENTS

1. Describe the series of reactions that eventually result in the blue color.
2. What is the "limiting" reagent in this reaction?
3. What effect does temperature have on reaction rate?
4. What effect does concentration have on reaction rate?
5. What other factors affect reaction rate?

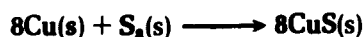
79. Catalytic Copper

Copper and sulfur powders are mixed and added to a hot blue solution of copper sulfate. A black solid forms and settles to the bottom of the flask. Repeating the procedure with hot water fails to form the black solid, copper sulfide.

Procedure

1. Bring 200 mL of copper sulfate solution to boiling in a 500-mL flask.
2. Mix 6.4 g of powdered copper and 3.2 g of powdered sulfur.
3. Stir into the copper sulfate solution.
4. Continue boiling the mixture, stirring, and swirling.

Reaction



Solutions and Materials

1. Copper sulfate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$: Dissolve about 1 tsp in 200 mL or around 0.1 M.
2. Copper and sulfur are powdered.

Teaching Tips

NOTES

1. Using twice as much copper as sulfur gives a *stoichiometric* mix. You might try an excess of sulfur to compare the yields visually.
2. It is important to stir constantly. Use a magnetic stirrer if you have one.
3. This reaction is a good example of a reaction between two solids.
4. Copper sulfate acts as a catalyst. The mechanism is not understood. Verify the catalytic effect of the solution by repeating the procedure using boiling water instead.
5. This demonstration can be a student experiment for conservation of mass. Have students design a procedure and suggest sources of error.

QUESTIONS FOR STUDENTS

1. Describe the differences that are apparent between the copper, sulfur, and copper sulfide.
2. Why is this a reaction between two solids?
3. Why do we use twice the mass of copper as sulfur?

80. Enzyme Kinetics: Effects of Temperature and an Inhibitor on Catalase Extracted from Potato

When catalase is added to hydrogen peroxide, a vigorous reaction occurs with the production of oxygen gas. When the demonstration is done at a higher temperature, the rate of oxygen production increases; when the demonstration is performed at a lower temperature, the rate decreases. When an inhibitor is added, enzyme activity decreases and the production of oxygen gas decreases.

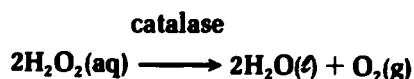
Effect of Temperature on Catalase Activity

Procedure

1. Prepare a gas-collecting apparatus by fitting a large test tube with a rubber tube leading from a one-hole rubber stopper. The other end of the tube should be directed beneath an inverted, water-filled, graduated cylinder in a large beaker or water trough filled with water.
2. Add about 5 mL of hydrogen peroxide, at room temperature, to the tube.
3. Quickly add 3 mL of catalase extract (see Solutions).
4. Stopper the test tube, swirl it once to mix, and collect the gas in the graduated cylinder.
5. Note the time required to collect 5 mL of the gas.
6. Repeat the demonstration, using both extract and hydrogen peroxide cooled in an ice bath. Record the temperature and the time required to produce 5 mL of oxygen.
7. Repeat, using extract and hydrogen peroxide warmed in a water bath to 40 °C. Again, record the temperature and the time required to collect 5 mL of oxygen.

Reaction

Catalase is an enzyme that specifically catalyzes the decomposition of hydrogen peroxide into water and oxygen gas.



The activity of an enzyme is increased with increased temperature and decreased with lowered temperatures. At a very high temperature, the enzyme (protein) is deactivated and no reaction occurs.

Solutions

1. Hydrogen peroxide, H_2O_2 , is the 3% drugstore variety.
2. Prepare the catalase extract as follows:
 - a. Peel and grind two to three Irish potatoes in a blender.
 - b. Add 25–30 mL of water to the mixture, swirl the mixture occasionally, and let it stand for about 10 min.
 - c. Filter the solution in cheesecloth to obtain a clear solution.
 - d. Use this solution as a source of catalase and polyphenyloxidase in this demonstration and demonstration 81.

Effect of an Inhibitor on Catalase Activity

Procedure

1. Place 5 mL of hydrogen peroxide in a test tube.
2. Place 5 mL of the catalase solution in a second test tube.
3. Add 5-10 drops of copper sulfate solution to the second test tube.
4. Place both test tubes in a water bath at 20 °C for 5 min.
5. Mix the test tubes and measure the amount of oxygen produced in 5 min.

Solutions

1. The copper sulfate concentration is 0.1 M: Dissolve 2.5 g of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in 100 mL of solution.
2. Hydrogen peroxide is a 3% solution.

Reactions

Catalase requires Fe^{2+} as an activator. Cu^{2+} ions compete with Fe^{2+} and prevent it from activating catalase.

Teaching Tips

NOTES

1. Most of the enzymes that occur in biochemical reactions have their maximum activity at body temperature (98.6 °F or 37.0 °C). However, some plant enzymes operate over a wider temperature range.
2. Catalase in the blood reacts with hydrogen peroxide when this solution is applied to an open wound. The bubbling effect is due to the production of oxygen gas.
3. Catalase is also present in rutabagas and turnips.
4. Notice that enzyme activity will not cease altogether, but it will decrease significantly.
5. Measure enzyme activity at various temperatures to find the temperature at which catalase has its maximum activity.

QUESTIONS FOR STUDENTS

1. What is an enzyme?
2. What does catalase do?
3. What effect does temperature have on enzyme activity?
4. How are enzymes inhibited? How can this inhibition be shown?

81. Enzyme Specificity: Polyphenoloxidase from Potato

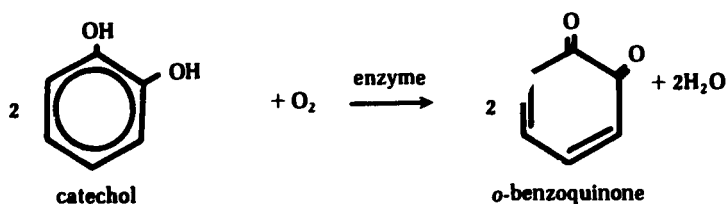
Enzymes act only on certain substrates. This fact is shown by adding polyphenoloxidase (from a potato extract) to catechol solution, which results in a red product. However, when the enzyme is added to substances with structures similar to that of catechol, no reaction occurs.

Procedure

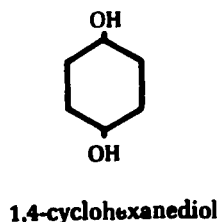
1. Prepare the enzyme extract from potato according to the directions given under Solutions in Demonstration 80.
2. Prepare and label four large test tubes as follows:
 - a. Test tube 1: 1 mL of distilled water.
 - b. Test tube 2: 1 mL of catechol solution.
 - c. Test tube 3: 1 mL of 1,4-cyclohexanediol solution.
 - d. Test tube 4: 1 mL of catechol solution.
3. Place all four test tubes in the water bath at 37 °C.
4. Measure 15 mL of the potato enzyme extract in another test tube. Place this test tube in the water bath also.
5. After 5 min, quickly pipet 3 mL of enzyme extract into each of the four test tubes, mix the solution quickly, and replace the test tubes in the water bath for an additional 5 min.
6. Observe the deep red color produced in the test tube containing the catechol but no color—or only a very light color—in the other test tubes.

Reactions

1. Polyphenoloxidase is a name for an enzyme that belongs to the oxidation-reduction group. This enzyme catalyzes the removal of hydrogen (oxidation) from dihydroxyphenols to molecular oxygen, and the corresponding quinones are formed.



2. All quinones are intensely colored, especially the o-quinones. Thus, the characteristic color of the products of this reaction can be used to measure the extent of enzyme activity.
3. Structures of the other compounds are



Solutions

1. The phenol solution concentration is 0.01 M. Be careful when using phenol. Use plastic gloves to handle the solid. If any is spilled on the skin, wash the skin immediately with soap and water.
2. The catechol concentration is 0.01 M: Dissolve 0.11 g of catechol per 100 mL of solution.
3. The 1,4-cyclohexanediol concentration is 0.01 M: Dissolve 0.12 g per 100 mL of solution. Use care with all of these reagents.

Teaching Tips

NOTES

1. The brown color that develops in bruised fruit and peeled potatoes is primarily due to this reaction.
2. All quinones are intensely colored, especially the *o*-quinones.
3. *Ortho* implies that the two oxygens on the benzene ring are next to each other.
4. This enzyme acts only on dihydroxyphenols. Although the other two compounds are similar, neither fit the requirements of this enzyme.
5. Distilled water is used as a standard.
6. Catechol is also called 1,2-dihydroxybenzene.

QUESTIONS FOR STUDENTS

1. What is meant by enzyme *specificity*?
2. Give some examples of common enzymes and their specific substrate. (Amylase and starch, catalase and hydrogen peroxide, and so forth.)
3. Why does this enzyme not react with phenol and 1,4-cyclohexanediol?
4. Propose a model for enzyme-substrate reaction.

82. Autocatalytic Effect

When a colorless solution is added to a purple solution of potassium permanganate, the purple solution becomes colorless in several minutes. However, when the process is repeated and a crystal of manganese chloride is added as a catalyst, the reaction occurs in about 30 s.

Procedure

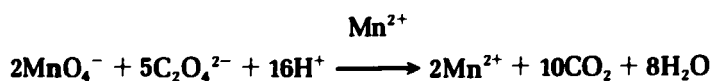
1. Prepare two 250-mL beakers as follows: 150 mL of solution A, 5 mL of concentrated sulfuric acid.
2. Add 50 mL of solution B to one beaker. Begin timing to see how long it takes for the purple color to completely fade. (Be patient! It may take several minutes.)
3. Add 50 mL of solution B to the second beaker, and add a few of crystals of manganese chloride.
4. Note the time required for the purple color in the second beaker to fade. (Adjust the amount for 20–30 s.)

Solutions

1. Solution A: oxalic acid. Add 8 g of oxalic acid to 500 mL of water. Stir until it dissolves.
2. Solution B: potassium permanganate, 0.001 M solution. Dissolve 0.16 g of KMnO_4 in 1 L.
3. Sulfuric acid is concentrated.

Reactions

1. Oxalate ions reduce the purple MnO_4^- ions to Mn^{2+} , which is a very pale pink that appears colorless in dilute solution. Mn^{2+} ions catalyze the reaction, but quite a bit of time is required for sufficient Mn^{2+} to build up; hence, the delay in the reaction.



2. In the second beaker, when Mn^{2+} is added, the reaction produces more Mn^{2+} , which in turn catalyzes the reaction; thus, the AUTOCATALYTIC effect is obtained.

Teaching Tips

NOTES

1. Vary the amounts until the reactions occur in a reasonable time.
2. Three to five minutes may be required for the first reaction to occur.
3. Any of these solutions may be disposed of by flushing them down the drain.

QUESTIONS FOR STUDENTS

1. What is the substance that is purple? What happens to it during the chemical reaction?
2. What is the catalyst for this reaction?
3. How much faster is the reaction in the second beaker? Why?

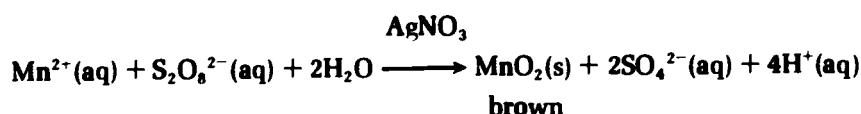
83. Oxidation of Manganese(II) Sulfate by a Catalyst

Manganese(II) sulfate plus a few crystals of another solid is placed in two test tubes and dissolved in water. To one test tube silver nitrate solution is added, and a brown precipitate is produced. The other test tube is heated, and a brown precipitate is produced.

Procedure

1. Place 0.5 g of manganese sulfate, MnSO_4 , into each of two test tubes and dissolve the MnSO_4 with 15 mL of water.
2. Add several crystals of potassium persulfate, $\text{K}_2\text{S}_2\text{O}_8$, to each test tube.
3. Add several drops of silver nitrate solution to the first test tube. Stir the contents of both tubes. Note the color change in the first test tube.
4. Gently heat the second test tube until a color change occurs.
5. Compare the color of the solutions in the two test tubes.

Reaction



Solution

The silver nitrate solution concentration is 0.2 M: Dissolve 3.5 g in 100 mL of solution.

Teaching Tips

NOTES

1. The brown precipitate formed is manganese dioxide.
2. The silver nitrate solution acts as a catalyst. Perhaps a silver complex is formed.
3. Heating causes the chemical reaction to take place without the catalyst.
4. Potassium persulfate, $\text{K}_2\text{S}_2\text{O}_8$, is also called potassium peroxydisulfate.

QUESTIONS FOR STUDENTS

1. What is the purpose of the silver nitrate solution?
2. Why is the second test tube heated?
3. What compound causes the brown color?
4. What is oxidized and what is reduced in this reaction?
5. Write the equation.

84. An Organic Clock Reaction

A light yellow aldehyde is mixed with a base and acetone. After about 20 s, a yellow precipitate suddenly appears. When the demonstration is repeated at a higher temperature, the precipitate forms in a shorter time.

Procedure

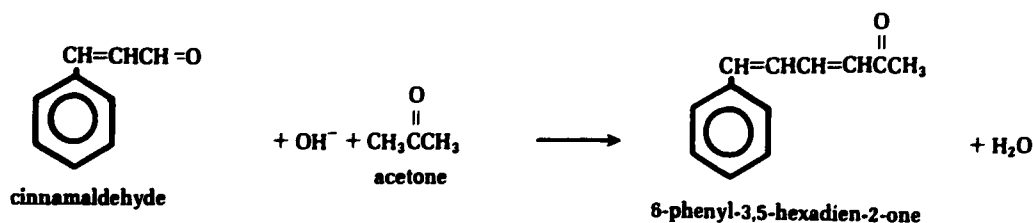
1. Place 50 mL of the aldehyde solution (see Solutions) in a 250-mL beaker. Use a tall form beaker, if one is available.
2. Add 30 mL of potassium hydroxide solution.
3. With constant stirring, add 2.0 mL of acetone.
4. Notice the time required for a yellow precipitate to form.
5. Repeat, using solutions at 40 °C.
6. Notice the time required for the precipitate to form.

Solutions

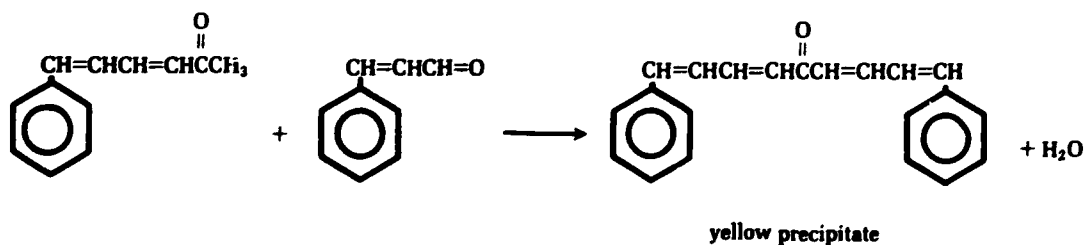
1. The aldehyde solution is made by dissolving 40 mL of cinnamaldehyde in 250 mL of 95% ethyl alcohol.
2. The potassium hydroxide solution concentration is 2.0 M: Dissolve 16.8 g of KOH in 150 mL of water.

Reactions

1. At least two reactions occur in this demonstration. The first reaction is a base-catalyzed aldol condensation between the aldehyde and the ketone (cinnamaldehyde and acetone).



2. When the acetone is consumed, the product from the first reaction can react with cinnamaldehyde to form the yellow precipitate.



Teaching Tips

NOTES

1. The yellow, insoluble product in the second reaction is 1,9-diphenyl-1,3,6,8-nonatetren-5-one.
2. Cinnamaldehyde is also 3-phenylpropenal.
3. Benzaldehyde can also be used. Add 32 mL to 250 mL of 95% ethyl alcohol. The reaction with benzaldehyde takes longer.
4. Use care with these organic reagents. Wash any spills on skin immediately.
5. Use ONLY a hot plate or water bath to heat the alcohol-aldehyde solution.
6. Because the reactions are temperature-dependent, the rate will approximately double for each 10-degree increase in temperature.
7. Cinnamaldehyde is an oil. It is the chief component of cinnamon oil.

QUESTIONS FOR STUDENTS

1. What is the role of potassium hydroxide in these reactions?
2. Which reaction, would you think, is the fastest?
3. What is the relationship of temperature to this reaction?

85. Equilibrium: The Dissociation of Acetic Acid

Several drops of indicator are added to acetic acid, and the solution turns light red. This solution is divided into two parts: Into one part, a small amount of solid sodium acetate is added, and the color of the solution changes to yellow. To the second part, a small amount of sodium hydroxide is added, and this solution, too, becomes yellow.

Procedure

1. Add 100 mL of acetic acid to a beaker.
2. Add a few drops of methyl orange indicator solution. Note the color of the acid solution.
3. Divide the solution into two smaller beakers.
 - a. Beaker A: Add a small pinch of solid sodium acetate. Note the appearance of a yellow (or tangerine) color.
 - b. Beaker B: Add sodium hydroxide dropwise until the color changes to yellow.
4. Ask students to explain this color change on the basis of equilibrium shifts

Reactions

This demonstration involves the equilibrium established when acetic acid dissociates in aqueous solution.



Methyl orange indicator is RED at pH 4.0 and YELLOW at pH 6.0.

1. Sodium acetate increases the concentration of CH_3COO^- on the right, which causes the equilibrium to shift to the left, reducing the H^+ concentration. The result is a yellow color.
2. Adding NaOH decreases the acidity, which forces the equilibrium to the right as the H^+ concentration is reduced. This reduction in hydrogen ion concentration produces a YELLOW color also.

Solutions

1. The acetic acid concentration is 1.0 M: See Appendix 2.
2. The sodium hydroxide concentration is 1.0 M. Add 40.0 g of NaOH per liter of solution.

Teaching Tips

NOTES

1. The yellow color may appear more tangerine in color.
2. Keep a sample of the original red solution for color comparison.
3. This is a good way to show the common ion effect.

QUESTIONS FOR STUDENTS

1. Explain how both sodium acetate and sodium hydroxide can reduce the hydrogen ion concentration.
2. Predict what would happen in each beaker if 1.0 M HCl were added. Try it!

86. Variations of the Formaldehyde Clock Reaction

Several beakers of solutions are set up, and different indicators are added to each beaker. After a few seconds, the solution in each beaker changes to a different color.

Procedure

1. Provide as many small beakers as you have indicators.
2. Add 10 mL of solution A to each beaker.
3. Add 10 mL of solution B to each beaker.
4. Place a few drops of each indicator in its labeled beaker. Note the original color in each beaker.
5. Add 10 mL of solution C to each beaker. After a few seconds, note the change in color of the solution in each beaker.

Reactions

This demonstration is a variation of the formaldehyde clock reaction. Essentially, OH^- produced in reaction 3 immediately reacts with HSO_3^- in reaction 4. When the HSO_3^- is consumed, OH^- causes a change in the indicator.

1. Reaction 1:



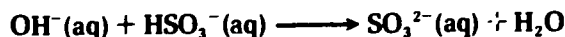
2. Reaction 2:



3. Reaction 3:



4. Reaction 4:



Solutions

1. Solution A: sodium sulfite. Prepare by adding 0.18 g of Na_2SO_3 to enough water to make 500 mL.
2. Solution B: sodium metabisulfite. Prepare by adding 0.79 g of sodium metabisulfite, $\text{Na}_2\text{S}_2\text{O}_5$, to enough water to make 500 mL.
3. Solution C: formalin. Dilute 5.5 mL of 37% formaldehyde with water to make 500 mL of solution.
4. Indicators: Any indicator that has a pH range of below 7.5 and above 9.0 will work. Try others.
 - a. Bromothymol blue: changes from yellow to blue. 0.1 g/16 mL of 0.01 M NaOH and 234 mL of water.
 - b. Cresol red: changes from yellow to red. 0.1 g/26 mL of 0.01 M NaOH and 224 mL of water.
 - c. o-Cresolphthalein: changes from colorless to red. 0.04% in ethyl alcohol.

- d. Metacresol purple: changes from yellow to purple. 0.1 g/26 mL of 0.01 M NaOH and 224 mL of water.
- e. Thymolphthalein: changes from colorless to blue. 0.04 g/50 mL of ethyl alcohol and 50 mL of water.
- f. Thymol blue: 0.1 g/22 mL of 0.01 M NaOH and 228 mL of water.

Teaching Tips

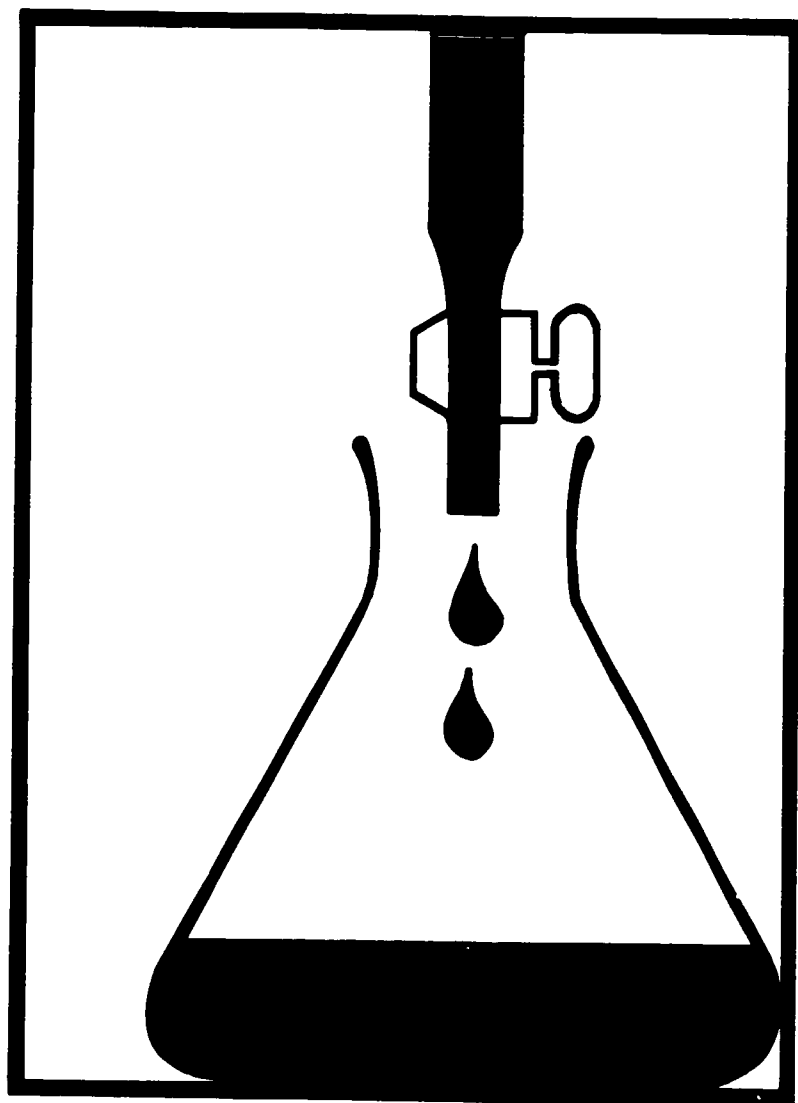
NOTES

1. A similar reaction, using phenolphthalein, is the basis for Demonstration 76.
2. Adjust the concentrations of solutions A–C and the amount of indicator to give the desired time for the clock to change colors.
3. Check the pH of the water—it could affect the reaction time.
4. Try 0.1 M $\text{Cd}(\text{NO}_3)_2$ as an indicator. It will change from colorless to white.
5. You can show the effect of temperature change on any of these clock reactions. Use reactants heated to various temperatures.
6. In the mechanism for the formaldehyde clock reaction, reactions 1 and 3 are slow and reactions 2 and 4 are fast.
7. Metabisulfite forms bisulfite in water solution.

QUESTIONS FOR STUDENTS

1. Explain the reactions that ultimately produce the color change in the clocks. Why do the colors eventually fade?
2. How can you account for the various colors?
3. Which of the reactions in the mechanism are slow and which are fast?
4. Which reaction is the rate-determining reaction?

Acids and Bases



87. Acid Rain

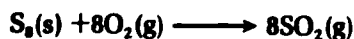
Sulfur dioxide gas, produced by burning sulfur, is dissolved in water containing universal indicator. The indicator changes color in the acid solution.

Procedure

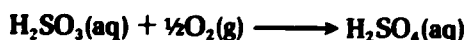
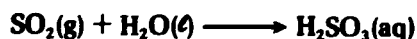
1. Prepare three large glass containers (about 1-L capacity) or flasks by adding about 100 mL of water to each.
2. Add enough universal indicator solution to each to give the water a distinct color. Usually green is the neutral range.
3. Set the first container aside as a standard.
4. To the second container add a few drops of acid to produce a color change in the solution. Set aside as an acid-color standard.
5. Carefully lower a deflagrating spoon containing a pea-size piece of burning sulfur into the third container. Use a FUME HOOD or provide adequate ventilation.
6. Note the production of a gas as the sulfur burns. Swirl the container and notice the color change as the gas dissolves in the water.

Reactions

1. Burning sulfur produces sulfur dioxide gas.



2. Sulfur dioxide dissolves in water to form sulfurous acid, $\text{H}_2\text{SO}_3(\text{aq})$, which oxidizes to form sulfuric acid.



The acid solution causes the universal indicator to change from green to red.

Solution

Acid: Any acid will work. Vinegar (5% acetic acid) works well.

Teaching Tips

NOTES

1. Sulfur dioxide is the primary air pollutant from the burning of coal by power plants. SO_2 dissolves in rain to produce the acid.
2. Sulfuric acid will react with marble (calcium carbonate) and many other substances to produce soluble calcium sulfate. This reaction results in the destruction of monuments, statues, gravestones, and so forth.
3. Ignite the sulfur by heating it carefully with a burner until it burns with a pale blue flame. The melting point of sulfur is 113°C .
4. Another simple method you can use to collect SO_2 is to heat the sulfur in a test tube, tilt the tube and ignite the sulfur, and collect the SO_2 gas in a large flask inverted over the test tube.

QUESTIONS FOR STUDENTS

1. How is acid rain produced?
2. What are the effects of acid rain? (Go to the library!)
3. How do you know that an acid was produced?

88. White Wine or Grape Juice?

Five beakers are arranged on the demonstration table. A pale yellow liquid (white wine) is poured from a large flask into the first beaker. When the liquid is poured into the second beaker, a blue liquid (grape juice) is produced. Both of these beakers are emptied into the large flask, and the entire contents turns blue. The blue liquid is poured into the third beaker, and a clear solution is produced. When the fourth beaker is filled, the liquid is blue again. When the last beaker is filled, the liquid is clear. When all the beakers are emptied back into the large flask, the entire contents of the flask becomes clear again.

Procedure

1. Arrange five 250-mL beakers, or clear glass tumblers, in full view of the class.
2. Prior to performing the demonstration, prepare the beakers as follows:
 - a. Beaker 1: empty.
 - b. Beaker 2: 2 drops of iron(III) chloride solution.
 - c. Beaker 3: 5 mL of sulfuric acid.
 - d. Beaker 4: empty.
 - e. Beaker 5: 5 mL of sulfuric acid.
3. Fill a 1-L flask, or large clear pitcher, with tannic acid solution.
4. Develop your own "story line": the following is a suggestion.
 - a. At a recent dinner party, a guest was served white wine. (Fill beaker 1 from the large flask, to produce "white wine" in the first beaker.)
 - b. The second guest asked for grape juice. Being a chemist, this is no problem. (Fill beaker 2 from the flask. A blue solution (grape juice) will be produced.)
 - c. If the guests prefer grape juice, then more will be needed. (Pour the contents of beakers 1 and 2 back into the flask. The entire contents of the flask will become blue.)
 - d. After making all this grape juice, the next guest asked for white wine! (Fill beaker 3 from the flask. A clear solution (white wine) is produced.)
 - e. The next guest asked for grape juice. (Fill beaker 4 from the flask. The liquid remains blue.)
 - f. But, the last guest asked for white wine. (Fill beaker 5 from the flask. A clear solution (white wine) is produced.)
 - g. Everyone wanted a second round of white wine! (Pour the contents of beaker 4 back into the flask and then pour the contents of beakers 3 and 5 back into the flask. The entire contents of the flask will become clear.)

Reactions

1. In beaker 2, tannic acid reacts with iron(III) chloride to form the blue complex iron(III) tannate.
2. When the blue iron(III) tannate is poured back into the flask, more of the blue complex is formed and the entire contents of the flask becomes colored.
3. When the blue "solution" is poured into beaker 3, sulfuric acid breaks up the iron(III) tannate complex and forms a colorless solution.
4. Pouring the contents of beakers 3 and 5 back into the flask breaks up all of the iron(III) tannate and produces a clear solution.

Solutions

1. Tannic acid: The concentration is not critical. About 1 tsp in 1 L of water works well. Adjust the amount of tannic acid until the solution is slightly yellowish.
2. Iron(III) chloride solution is saturated.
3. Sulfuric acid is concentrated. CAREFUL WITH THIS!

Teaching Tips

NOTES

1. The formula for tannic acid is $C_{76}H_{52}O_{46}$.
2. Adjust the concentration of iron(III) chloride until a deep blue color is produced when the tannic acid solution is added.
3. The chemical reactions involved in this demonstration are not well-defined.

QUESTIONS FOR STUDENTS

1. Why does the colorless solution become blue in color?
2. Why does the blue solution become colorless again?
3. Can you think of a use for iron tannate? [Iron(II) tannate is used to make ink. It is easily oxidized in air to form the complex iron(III) tannate.]

89. Amphoteric Properties of Metal Hydroxides

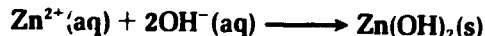
A clear solution of sodium hydroxide is added to a clear solution of zinc chloride. A white precipitate forms and is divided into two parts. As more sodium hydroxide is added to one part, the white precipitate dissolves. When an acid is added to the second part, the precipitate dissolves also.

Procedure

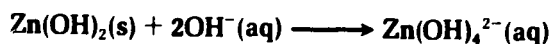
1. Place 100 mL of zinc chloride solution in a large beaker.
2. Place 250 mL of sodium hydroxide solution in a second beaker.
3. Place 250 mL of hydrochloric acid in a third beaker.
4. Slowly add the sodium hydroxide solution to the zinc chloride solution.
5. Note the formation of a dense, white, gelatinous precipitate of zinc hydroxide.
6. Pour about half of the precipitate into another beaker.
7. Continue to add sodium hydroxide solution to the precipitate in one beaker.
8. Note that the white precipitate dissolves.
9. Add hydrochloric acid to the precipitate in the other beaker.
10. Note that the precipitate also dissolves in acid.

Reactions

1. When sodium hydroxide (or any strong base) is added to the zinc chloride solution, the white precipitate zinc hydroxide forms.



2. With additional base, zinc hydroxide forms the complex ion $\text{Zn}(\text{OH})_4^{2-}$, which is soluble.



3. The metal hydroxide dissolves in hydrochloric acid (or any strong acid).



Solutions

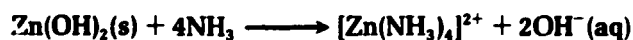
1. The zinc chloride solution concentration is 0.5 M: Dissolve 68 g of ZnCl_2 in 1 L of solution.
2. The sodium hydroxide solution concentration is 1.0 M: Dissolve 40 g of NaOH in 1 L of solution.
3. The hydrochloric acid concentration is 1.0 M: See Appendix 2 for dilution.

Teaching Tips

NOTES

1. An amphoteric hydroxide reacts with acids as well as with bases.
2. Zinc hydroxide, like most metal hydroxides except those of group 1 and 2 elements, is insoluble.
3. Aluminum, tin, lead, and chromium also form amphoteric hydroxides.

4. $[\text{Zn}(\text{OH})_4]^{2-}$ is called the *zincate ion* or tetrahydroxozincate(II).
5. The solubility product constant, K_{sp} , for $\text{Zn}(\text{OH})_2$ is 4.5×10^{-17} .
6. Zinc hydroxide is also soluble in aqueous ammonia, forming tetraamminezinc hydroxide. Try it!



7. The concentrations of solutions in this demonstration are not critical. Try varying the concentrations to produce the desired results.

QUESTIONS FOR STUDENTS

1. What is meant by the term *amphoteric*?
2. Explain how zinc hydroxide can be amphoteric.
3. Name the product formed when zinc hydroxide reacts with an acid and with a base.
4. What other metal hydroxides are amphoteric?

90. Milk of Magnesia versus Acid

A Milk of Magnesia tablet dissolves and reacts with dilute acid in the presence of universal indicator. Several color changes result as this preparation for stomach disorder reacts.

Procedure

1. In a 500-mL beaker, add 300 mL of distilled water and place the beaker on a magnetic stirrer.
2. Add 6 drops of 1.0 M HCl and 24 drops of universal indicator.
3. Drop in one tablet of Phillips Milk of Magnesia.
4. Observe the color changes.

Reactions

The active ingredient in the tablet is magnesium hydroxide, Mg(OH)_2 . The solution becomes neutral when it reacts with the acid and the excess is dissolved to form the eventual basic solution.



Solution

The concentration of the hydrochloric acid is 1.0 M: See Appendix 2.

Teaching Tips

NOTES

1. This reaction takes up to half the class period and is one that the students can be watching during discussion or lecture.
2. The solution starts out about pH 4, and the indicator will be red.
3. The color changes are red to orange to yellow (pH 6.5) to green to blue (pH 9) to violet to red-violet (pH 10).
4. No buffering action occurs with magnesium hydroxide.
5. Try Tums (calcium carbonate) or Rolaids (dihydroxylaluminum sodium carbonate). Because a buffer forms in each of these reactions, the pH is maintained at 6.5 when all the tablet has dissolved.
6. Although magnesium hydroxide is a strong base, it will cease to raise the pH as it reaches its point of solubility—0.0009 g/100 mL at 18 °C.

QUESTIONS FOR STUDENTS

1. Milk of Magnesia is an antacid. How does an antacid work?
2. Why did the pH change in the beaker?
3. Why will the pH not go over 10.0 even though more Milk of Magnesia is in the beaker?

91. Simple Buffer Action

Two beakers are displayed. One beaker contains water and the other beaker contains a buffer solution; both contain an indicator and are at pH 7. When a base is added to the water, the water changes color immediately; but, several milliliters are required to change the color of the indicator in the buffer. The process is repeated and shows the effect of adding acid to water and a buffer solution.

Procedure

1. Place 100 mL of water in a 250-mL beaker.
2. Place 100 mL of buffer solution (see Solutions) in another beaker.
3. Add 5–8 drops of phenolphthalein indicator solution to each beaker.
4. With the assistance of two students, add sodium hydroxide solution, dropwise, to each beaker.
5. Record the number of drops required to obtain a permanent pink color in each beaker.
6. Repeat steps 1 and 2. Add the same number of drops of methyi orange indicator solution to each beaker.
7. Again with the help of two students, add hydrochloric acid dropwise until a red color is produced.
8. Interpret these results in terms of buffer action.

Reactions

A buffer solution resists change in pH when either a base or an acid is added. A buffer solution consists of a weak acid and the anion that is its conjugate base.

The phosphate buffer can react with acids or bases as follows:

1. $\text{HPO}_4^{2-}(\text{aq}) + \text{H}^+(\text{aq}) \longrightarrow \text{H}_2\text{PO}_4^-(\text{aq})$
2. $\text{H}_2\text{PO}_4^-(\text{aq}) + \text{OH}^-(\text{aq}) \longrightarrow \text{HPO}_4^{2-}(\text{aq}) + \text{H}_2\text{O}(\ell)$

Solutions

1. The buffer solution: Prepare the pH 7.0 buffer solution by dissolving 7.0 g of potassium acid phosphate, $\text{K}_2\text{H}_2\text{PO}_4$, in 295 mL of 0.1 M KOH and diluting to 1 L.
2. The sodium hydroxide solution concentration is 0.1 M: Dissolve 4.0 g of NaOH in 1 L of solution.
3. The hydrochloric acid solution concentration is 0.1 M: See Appendix 2.

Teaching Tips

NOTES

1. To show the effect of adding acid to the buffer solution, you could also use bromocresol green to get a blue color or bromocresol blue to get a yellow color. Try other indicators.
2. Buffers are important in maintaining a plasma pH of 7.4. The phosphate buffer accounts for about 10% of the total blood buffers.

3. Essentially, buffers contain a base to react with added acid and an acid to react with added base.
4. Try other buffers. The acetate-acetic acid buffer is easy to prepare and to work with.

QUESTIONS FOR STUDENTS

1. What is a *buffer*?
2. How can a buffer accept either an acid or a base without changing pH?
3. How does the buffer in this demonstration work?
4. What are some practical applications of buffers?

92. Disappearing Ink

A blue liquid is made in the laboratory. When this liquid is squirted on a white shirt or dress, it leaves a blue "ink" spot. In a few seconds, the blue spot fades and eventually disappears!

Procedure

1. Place about 50 mL of ethyl alcohol in a beaker.
2. Add a few drops of thymolphthalein indicator solution.
3. Add just enough NaOH solution, dropwise, to produce a deep blue color in the solution.
4. Put the "disappearing ink" in a small squirt bottle.
5. Enjoy!

Reactions

1. Sodium hydroxide reacts with carbon dioxide in the air to form sodium carbonate, which is less basic than sodium hydroxide.



2. The less basic sodium carbonate causes the indicator to change from BLUE to COLORLESS. The colorless range for thymolphthalein is below pH 9.3. The blue range is above pH 10.5, and the color change takes place between pH 9.3 and pH 10.5.
3. The alcohol evaporates and leaves only the colorless residue.

Solutions

1. The ethyl alcohol concentration is 95%.
2. NaOH is a 1.0 M solution (concentration is not critical).
3. For thymolphthalein, use either small amounts of solid or a solution in 50% water-50% ethyl alcohol.

Teaching Tips

NOTES

1. This ink is the same "disappearing ink" as that sold in trick shops and novelty stores.
2. Adjust the amount of indicator to get a deep blue solution.
3. The compound produced, Na_2CO_3 , is actually *washing soda*.

QUESTIONS FOR STUDENTS

1. Why does the ink evaporate so quickly?
2. Show how this is an "acid-base" reaction.
3. Why is thymolphthalein a good indicator for this reaction?

93. Colorful Effects of Hydrochloric Acid Dilution

Indicators are added to decreasing concentrations of HCl with colorful results.

Procedure

1. Prepare 100 mL of 1.0 M HCl and pour into beaker 0, labeled pH 0.
2. Remove 10 mL of the 1.0 M HCl, dilute to 100 mL, and pour into beaker 1, labeled pH 1.
3. Remove 10 mL of the 0.1 M HCl solution from beaker 1, dilute to 100 mL, and pour into beaker 2, labeled pH 2.
4. Remove 10 mL of the 0.01 M HCl solution from beaker 2, dilute to 100 mL, and pour into beaker 3, labeled pH 3.
5. Continue the serial dilutions to prepare beakers with pH 4 and 5.
6. Divide each beaker to form two samples of each pH. Arrange in pairs.
7. Add several drops of methyl violet indicator to each pH in one set.
8. Add several drops of methyl orange indicator to each pH of the other set.
9. Identify the appropriate indicator for each pH value.

Reaction

An indicator (In) is a weak acid in equilibrium with its conjugate base.



When the hydrogen ion concentration is increased, the equilibrium shifts to the left and the color is in its "acid" form. Different indicator molecules respond to different hydrogen ion concentrations.

Solution

The concentration of HCl is 1.0 M. See Appendix 2.

Teaching Tips

NOTES

1. This demonstration works well in small beakers or plastic cups on the overhead projector. Solution volumes must be the same.
2. Colors of methyl violet range from yellow-green to turquoise to blue and violet. Methyl orange is red up to pH 4 and changes to tangerine yellow above pH 4.2.
3. Additional sets may be prepared and other indicators used.

Indicator	pH Range
Thymol blue	1.2-2.8
Bromophenol blue	3.0-4.6
Methyl red	4.6-6.3

4. The methyl violet is the only indicator that gives four colors.
5. Try preparing natural indicators from tea, red cabbage, red roses, blueberries, and other flowers.
6. The number of dilutions compound the error so that at pH 6 the concentration is probably not accurate.

QUESTIONS FOR STUDENTS

1. What is an *indicator*, and how does it work?
2. What is the concentration of H^+ and the pH in each beaker?
3. Which indicator is most useful for identifying each pH?

94. The Boiler-Scale Reaction

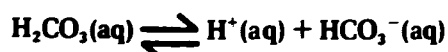
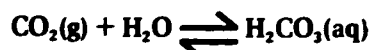
Dry ice (solid carbon dioxide) is added to a flask of clear liquid. A precipitate forms, redissolves, and re-forms when the flask is heated.

Procedure

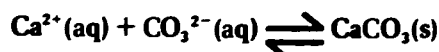
1. Place 300 mL of clear limewater (calcium hydroxide solution) in a 500-mL flask.
2. Drop in a chunk of dry ice.
3. Observe the formation of a precipitate and the redissolving to form a clear solution again.
4. Heat over a low flame or on a hot plate. Observe the reprecipitation.

Reactions

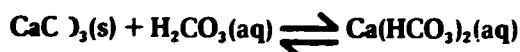
1. Carbon dioxide reacts with water to form CO_3^{2-} ions as one of the products.



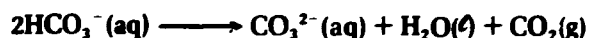
2. The carbonate ions react with the calcium ions to form insoluble calcium carbonate.



3. The additional dissolving of carbon dioxide and formation of acid cause the calcium carbonate to dissolve.



4. Upon heating, the calcium hydrogen carbonate decomposes.



The calcium carbonate reprecipitates as shown in reaction 2.

Solutions and Materials

1. Limewater is dilute calcium hydroxide solution. Add a few teaspoonfuls of calcium hydroxide or calcium oxide to 1 L of water. Prepare the day before the demonstration to allow undissolved solid to settle.
2. Dry ice is solid carbon dioxide.

Teaching Tips

NOTES

1. Ice cream shops and truck stops are good sources of dry ice.

2. This reaction series regularly takes place in hard water. When the water is heated, calcium, magnesium, and iron carbonates precipitate to form *boiler scale*.
3. When the carbonates precipitate in heated water, the water is effectively "softened". This produces temporary hard water. Permanent hard water has anions other than carbonates and cannot be softened by heating.
4. The acid nature of solutions of carbon dioxide accounts for the formation of caves. Reprecipitation of calcium carbonate forms stalactites and stalagmites.

QUESTIONS FOR STUDENTS

1. What causes the initial precipitation?
2. What causes the redissolving of the precipitate?
3. Why does the precipitate re-form when the solution is heated?
4. Why would boiler scale be a particular problem in a steam power plant? (The scale acts as an insulator, which may cause the boiler to overheat, which presents an explosion danger.)

95. Reversible Oxidation-Reduction Color Changes

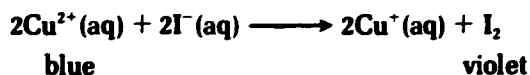
Reversible color changes are seen in two immiscible liquids in a large test tube as a result of shifts in the equilibrium. When colorless iodide solution is added to blue copper(II) solution, a green solution results and purple molecular iodine appears in the nonaqueous layer. Additional color changes result when ammonia is added, then neutralized with acid.

Procedure

1. In a large test tube, place about 5 mL of trichlorotrifluoroethane or other nonpolar solvent with a density greater than 1.0 g/cm³.
2. Add about 25 mL of 0.05 M copper sulfate solution.
3. Stopper and shake. Observe the separation of the layers of two liquids and the blue color of the aquated copper(II) ions.
4. Add a small amount, 1-2 mL, of 0.2 M KI.
5. Stopper and shake, observing the presence of the violet molecular iodine in the layer of nonpolar solvent.
6. Add concentrated ammonia dropwise to the tube until a clear deep blue solution results. Stopper and shake.
7. Observe the resulting color changes, including the decolorization of the nonaqueous layer.
8. Add 8 M HNO₃ dropwise to neutralize the ammonia and observe the shift toward formation of violet iodine again.

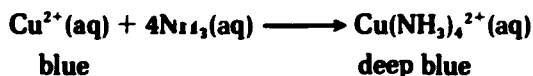
Reactions

1. When copper ions and iodide ions react, the copper is reduced to copper(I) and the iodide is oxidized to form molecular iodine, I₂, which is violet in the organic solvent layer.



The copper(I) ions form a soluble complex in the potassium iodide solution.

2. The added ammonia complexes with the Cu²⁺(aq) to form Cu(NH₃)₄²⁺(aq), tetraamminecopper(II), thus removing it from the reaction with iodide.



The equilibrium of reaction 1 shifts as Cu²⁺ is removed. Reaction 1 proceeds towards the left to compensate for the lowered Cu²⁺ concentration. The molecular iodine concentration is then reduced, and the layer loses its violet color. Adding acid to the system removes the ammonia, and thus increases the Cu²⁺ concentration. Reaction 1 proceeds to form molecular iodine again.

Solutions

1. Any standard nonpolar solvent will work if it has a density greater than water. Trichlorotrifluoroethane is a common substitute for carbon tetrachloride.
2. Copper sulfate, CuSO₄·5H₂O, concentration is 0.05 M: 12.5 g/L.

3. Potassium iodide, KI, concentration is 0.2 M: 3 g/100 mL.
4. Ammonia is concentrated.
5. Nitric acid is 8 M: See Appendix 2.

Teaching Tips

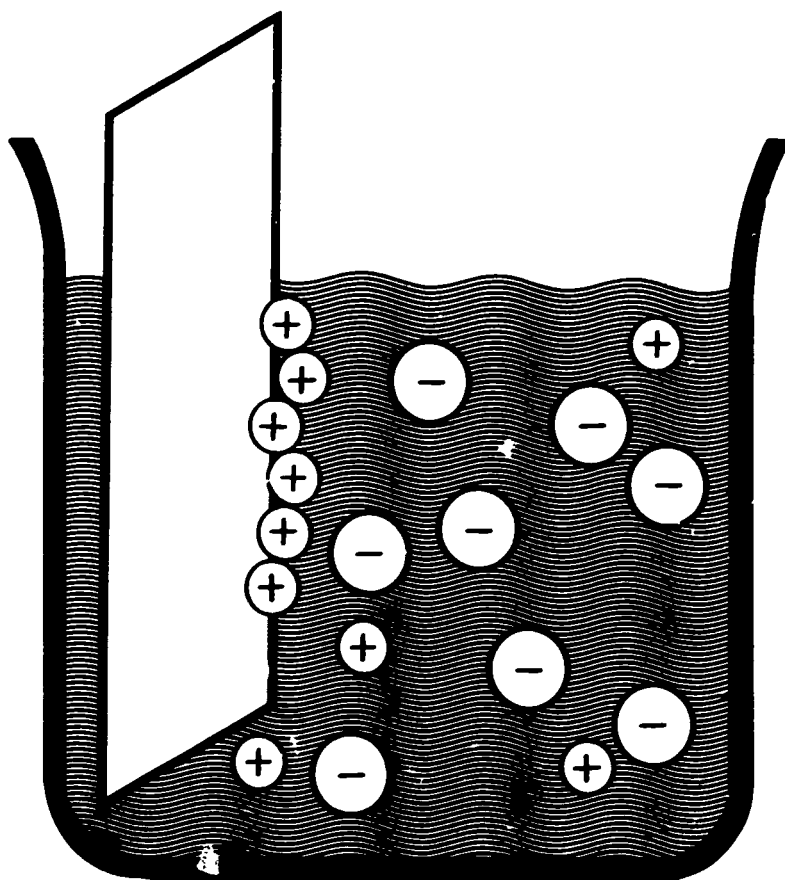
NOTES

1. A number of principles are reviewed in this demonstration in addition to equilibrium: solubilities, densities of liquids, colors of copper complexes and iodine, and oxidation and reduction.
2. The concentrations of the solutions are not critical.
3. The copper(I) in aqueous solution is not stable and will eventually disproportionate to form copper(II) and copper(0).
4. The greenish solution that forms upon addition of potassium iodide is a mixture of soluble complex ions of copper(I) and copper(II). If the copper and iodide are more concentrated than given here, precipitation of CuI (white) will be seen as well as the formation of brown triiodide ion.
5. This reaction system is also easily viewed in a separatory funnel.

QUESTIONS FOR STUDENTS

1. How can you identify the presence of aquated copper(II) in solution?
2. How can you identify the presence of tetraamminecopper(II) in solution?
3. How can you identify the presence of molecular iodine?
4. What caused the shift in equilibrium? What observations caused you to identify the shift?

Oxidation and Reduction Reactions



96. Changing Colors: Orange to Green

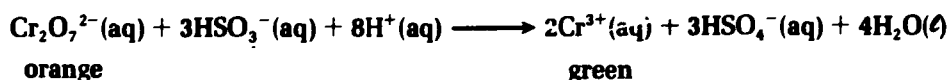
A folded white paper napkin is dipped into an ORANGE solution. When the napkin is blotted and dipped into a colorless solution, the napkin turns GREEN.

Procedure

1. Place about 100 mL of potassium dichromate solution in a large beaker.
2. Place the same amount of sodium hydrogen sulfite solution in a second beaker.
3. Fold a white paper napkin, or handkerchief, several times. Carefully dip the napkin into the potassium dichromate solution. Note the ORANGE color of the napkin.
4. Blot the excess liquid on the napkin and dip it into the sodium hydrogen sulfite solution. Notice the GREEN color of the napkin.

Reaction

The color change is due to the formation of the green Cr^{3+} ion, produced when bisulfite ion reacts with the dichromate ion.



Solutions

1. The concentrations of the solutions are not critical to the demonstration.
2. The following work well:
 - a. Potassium dichromate: Dissolve 1.0 g of $\text{K}_2\text{Cr}_2\text{O}_7$ in 100 mL of water.
 - b. Sodium hydrogen sulfite: Dissolve 5 g of NaHSO_3 in 100 mL of water.

Teaching Tips

NOTES

1. This demonstration is a colorful way to show oxidation-reduction.
2. At the end of the demonstration, pour the remaining sodium hydrogen sulfite solution into the dichromate solution—it will all turn green.
3. Sodium hydrogen sulfite is also called sodium bisulfite.

QUESTIONS FOR STUDENTS

1. Why did the napkin turn green?
2. Which chromium ion has a characteristic green color? What colors characterize other chromium ions? (Look it up!)
3. What is oxidized and what is reduced in this reaction?

97. The Copper Mirror

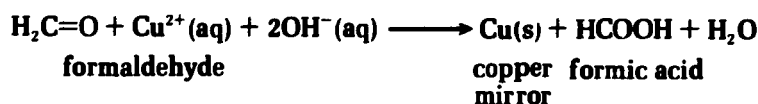
Two solutions mixed in a test tube and heated form a shiny mirror of copper metal coating the inside walls of the test tube.

Procedure

1. Stuff a wad of 000 steel wool into the bottom of a 10-mL test tube and, using a forceps, thoroughly roughen the inside of the tube.
2. Wash the inside of the tube with concentrated nitric acid. Rinse well.
3. Rinse with a small amount of tin(II) chloride solution.
4. Rinse with a small amount of silver nitrate solution.
5. Add 10 mL of solution A to 10 mL of solution B.
6. Place 3 mL of this mixture in the test tube.
7. Add 3 mL of solution C.
8. Place the test tube in a hot water bath or heat over a direct flame.
9. The copper mirror should develop in a few minutes.

Reaction

The formaldehyde is oxidized to formic acid, and the copper is reduced. In the typical reaction using these reagents, the copper is reduced to copper(I). However, because the formaldehyde is quite concentrated, further reduction takes place to form metallic copper.



Solutions

1. Solution A: Fehlings reagent I. Dissolve 3.5 g of copper sulfate in 50 mL of water.
2. Solution B: Fehlings reagent II. Dissolve 3 g of sodium potassium tartrate and 6.0 g of sodium hydroxide in 50 mL of water.
3. Solution C: Formalin. 37% formaldehyde (must be fresh).
4. Tin(II) chloride is a 1% solution.
5. Silver nitrate is a 1% solution.

Teaching Tips

NOTES

1. The fine scratching and rinsing of the tube are necessary for good results.
2. If you suspect your mirror to be copper(I) oxide, try a small amount of hydrochloric acid to dissolve it.
3. Heating of the solutions should be kept under the boiling point.
4. Because formaldehyde gas may be evolved, you may wish to heat the mixture under the hood.
5. After the copper mirror forms, flush the solutions down the drain and rinse the test tube.

QUESTIONS FOR STUDENTS

1. Find a procedure for the formation of the silver mirror. Compare the reactions.
2. In the formation of the copper mirror, what is being oxidized and what is being reduced?
3. Write the structural formulas for formaldehyde and formic acid.

98. Reduction of Copper Oxide

A test tube containing copper oxide and carbon is heated. A gas and metallic copper are produced. The copper forms a coating on the side of the tube.

Procedure

1. Place about $\frac{1}{4}$ tsp of copper(II) oxide in a test tube.
2. Add $\frac{1}{2}$ tsp of powdered charcoal (carbon).
3. Mix the two chemicals by shaking the tube.
4. Stopper the tube with a one-hole rubber stopper with a glass delivery tube.
5. Place the delivery tube in a second test tube half filled with limewater.
6. Heat the mixture in the tube, gently at first and then strongly.
7. Notice the evolution of a gas and the formation of a white, milky precipitate in the tube containing limewater.
8. After the reaction has ceased, cool the tube and observe the copper on the inside of the tube.

Reactions

1. This demonstration is an oxidation-reduction reaction. Carbon reduces the copper oxide to form copper. Carbon is then oxidized to carbon dioxide.



2. Carbon dioxide reacts with limewater (calcium hydroxide) to form a precipitate of calcium carbonate. This reaction is the traditional test for carbon dioxide.



Solution

Limewater: a saturated solution of calcium hydroxide. Add calcium hydroxide until no more will dissolve. Let the solution stand overnight. Decant the clear solution.

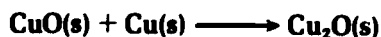
Teaching Tips

NOTES

1. Try varying the amount of chemicals and the size of the test tube to get the best yield of copper.
2. Copper(II) oxide can be made by strongly heating copper(I) oxide in oxygen.



3. If CuO is heated strongly with copper powder, copper(I) oxide is formed.



QUESTIONS FOR STUDENTS

1. What is the role of carbon in this demonstration?
2. What evidence exists that carbon dioxide is produced?
3. What is the coating in the test tube?

99. Corrosion of an Iron Nail

An iron nail is placed in a prepared gelatin solution. Within an hour or so, the areas around the head and the tip of the nail have turned blue and the area around the body of the nail has turned pink.

Procedure

1. Prepare a Petri dish with gelatin as indicated under Solution.
2. Place a nail in the dish containing the gelatin and allow the gelatin to become firm.
3. Notice, after an hour or so, that the areas around the head and the tip of the nail have become blue, whereas the area around the body of the nail has become pink.

Reactions

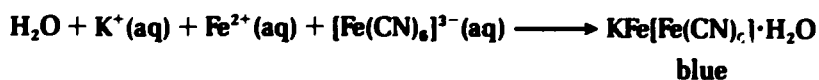
This demonstration is an oxidation-reduction reaction.

1. Oxidation: Electrons are given up at the points of the nails because the strained metal is more anodic (more active) than the body of the nail. Iron is oxidized to Fe^{2+} .

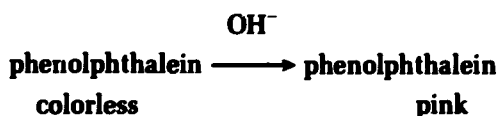
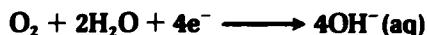


Fe^{2+} reacts with $\text{K}_3[\text{Fe}(\text{CN})_6]$ in two steps to produce a blue precipitate known as *Turnbull's blue*.

Fe^{2+} reacts with $\text{K}_3[\text{Fe}(\text{CN})_6]$ to produce a blue precipitate.



2. Reduction: Electrons given up at the two ends pass along the body of the nail where they are used to reduce oxygen dissolved in the gel to OH^{-} . The increased concentration of OH^{-} causes phenolphthalein to change to a pink color.



Solution

Prepare the gelatin for the Petri dish by adding a few drops of phenolphthalein solution and about 1 mL of 0.1 M potassium ferricyanide solution (33 g/L of solution) to 250 mL of warm water into which a package of clear gelatin has been dissolved.

Teaching Tips

NOTES

1. This demonstration is similar to "Red and Blue Electron Trail" (Demonstration 101).
2. You can use agar instead of gelatin.
3. This demonstration projects well on an overhead projector.
4. You can easily expand upon this demonstration.
 - a. Bend the nail to produce another strained area (the bend). The area near the bend will become blue also.
 - b. To show that a more active metal (like zinc) will protect iron from corroding wrap a piece of zinc foil or metal around the middle of the nail and repeat the demonstration. Zinc more readily loses electrons than iron so no Fe^{2+} is produced to form "rust". No blue areas will develop around the nail.
 - c. Try magnesium ribbon wrapped around the nail.
5. Copper or aluminum would not protect iron from corroding, because they are less easily oxidized than iron.
6. Aluminum corrodes slightly to produce Al_2O_3 . This film covers the metal and prevents further corrosion. This slight corrosion is why aluminum products (like cans) have a long lifetime.
7. If you use galvanized nails, dip them in dilute HCl to remove the zinc coating.

QUESTIONS FOR STUDENTS

1. Explain the oxidation and reduction reactions in this demonstration.
2. Why are parts of the nail surrounded by a blue color and other parts by a pink color?
3. How can you prevent the corrosion (blue color) from occurring?
4. Why do aluminum cans last so long, whereas iron cans gradually rust away?

100. The Mercury Amoeba

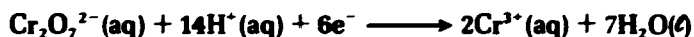
A pool of mercury in a Petri dish flattens and flows into a variety of shapes when a small amount of solution is applied.

Procedure

1. Place a small pool (~0.5 cm in diameter) of mercury in a Petri dish.
2. Add enough nitric acid to cover the bottom of the Petri dish.
3. Using forceps, place a few large crystals of solid potassium dichromate, $K_2Cr_2O_7$, on the dish, touching the pool of mercury.
4. Project the "amoeboid" movement of the mercury on an overhead projector.

Reactions

Oxidation of the mercury by the dichromate changes the charge on the surface. This charge results in the movement of the pool of mercury and the change in its shape.



Solution

The nitric acid concentration is 1.0 M: See Appendix 2 for dilution.

Teaching Tips

NOTES

1. The mercury need not be clean.
2. Be careful with mercury. Spills should be cleaned up immediately.
3. You may also observe a red precipitate of mercury(II) dichromate. If the amount of precipitate obscures the view of the amoeba, remove the dish from the projector and observe the dish directly.
4. Wash the mercury after use and store it in a bottle for reuse.
5. Important concepts in this demonstration include surface tension, electron transfer, and oxidation-reduction.

QUESTIONS FOR STUDENTS

1. Why does the "amoeba" move?
2. What is oxidized and what is reduced?
3. What is the red precipitate formed in the reaction?
4. What factors would determine how long the "amoeba" could move?

101. A Red and Blue Electron Trail

Two iron nails are attached with wire leads to the terminals of a 6-V cell. The nails are placed in a beaker of gelatin. After a minute or so, the area around one nail turns BLUE, and the area around the other nail turns PINK. The flow of electrons can be traced from the cell through the pink area, then through the blue area, and back to the cell.

Procedure

1. Prepare the beaker of special gelatin (see Solutions).
2. Obtain a 6-V cell.
3. Attach wire leads from the cell to two clean iron nails.
4. Stick the nails in the beaker of gelatin. Separate the nails as much as possible.
5. Observe. For best results, leave the cell attached for about 1 h.

Reactions

AT THE CATHODE

1. Electrons from the nail cause the reduction of water.



2. Phenolphthalein reacts with the OH^- to form the PINK color.

AT THE ANODE

1. Oxidation of the other nail produces electrons, which enter the circuit.



2. Fe^{2+} reacts with $\text{K}_3\text{Fe}(\text{CN})_6$ to produce the BLUE complex.



Solutions

Prepare several beakers as follows:

1. Follow the directions on a commercial gelatin box to make several hundred milliliters.
2. Add 10 drops of phenolphthalein solution. Phenolphthalein solution is made by dissolving 1.0 g of phenolphthalein in 50 mL of ethyl alcohol and 50 mL of water.
3. Add 10 mL of potassium ferricyanide, $\text{K}_3\text{Fe}(\text{CN})_6$ solution. The $\text{K}_3\text{Fe}(\text{CN})_6$ solution concentration is 0.1 M: Dissolve 33 g in 1 L of water.
4. Fill several beakers with the solution and allow the solution to cool and solidify.

Teaching Tips

NOTES

1. The BLUE color is produced at the ANODE.
2. The PINK color is produced at the CATHODE.
3. The blue complex, $\text{Fe}_3[\text{Fe}(\text{CN})_6]_2$, is also called *Turnbull's blue*.
4. The nail at the cathode could never rust. Why? (The cathode is supplying electrons; hence, no iron oxide could form on the nail. This property is called *cathodic protection*.)
5. You may want to use two 6-V cells to speed up the reaction.
6. Potassium ferricyanide is also called potassium hexacyanoferrate(III).

QUESTIONS FOR STUDENTS

1. What substance must be present for the pink color to appear? How is this substance produced?
2. The blue color is from iron(II) ions reacting with the yellow substance, potassium ferricyanide. How are the Fe^{2+} ions formed?
3. Trace the flow of electrons from the cell into the beaker-nail assembly and back to the cell.

102. Reduction of Sand with Magnesium

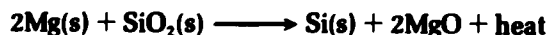
A mixture of sand (silica, SiO_2) and magnesium powder is heated in a test tube. As a result of this highly exothermic reaction, the brightly glowing mixture produces powdered magnesium oxide and coats the tube with a silicon mirror.

Procedure

1. Obtain a large Pyrex test tube (You will not be able to reuse the test tube after the demonstration).
2. Mix 5 g of magnesium powder and 7.5 g of dry silica sand and place this mixture in the dry test tube.
3. Clamp the tube in position on a ring stand.
4. Heat the bottom of the tube until an orange glow indicates that the reaction has begun.
5. Remove the heat and notice that the reaction continues.
6. Let the tube cool. Notice the formation of a shiny gray silicon mirror and a deposit above the mirror of white magnesium oxide.

Reaction

Magnesium metal reduces silicon dioxide to form silicon and magnesium oxide.



Materials

1. Magnesium powder: Fine powder works best.
2. Silica sand.

Teaching Tips

NOTES

1. This reaction is a very exothermic and should be done behind a safety shield.
2. The heat of the reaction will probably partially melt the test tube.
3. Silicon dioxide occurs in three crystalline forms. All have a SiO_4 tetrahedral structure.
4. Silicon is the second most abundant element in the Earth's crust, but it is not found naturally as the free element.
5. Silicon is made commercially by the reduction of sand and is used in steel, iron, and aluminum alloys and as the starting material for the production of silicone polymers.
6. Silicon was first prepared by Berzelius in 1823 by heating potassium and silicon tetrafluoride.

QUESTIONS FOR STUDENTS

1. What is oxidized in this reaction? What is reduced?
2. How does the silicon mirror form?
3. How is silicon prepared commercially?
4. What are some uses for silicon?
5. How does SiO_2 compare with CO_2 ? (Look it up!)

103. Electroplating Copper

A copper strip and a stainless steel spoon are immersed in a solution of copper sulfate. Each is connected to a 1.5-V battery with wire leads. After a few minutes, the spoon is coated with red-brown copper.

Procedure

1. Clean a copper strip by dipping it into a beaker of dilute nitric acid and then washing it thoroughly. CAREFUL with nitric acid!
2. Attach the copper strip to the POSITIVE terminal of a 1.5-V battery using a wire lead with an alligator clip.
3. Attach the stainless steel spoon to the NEGATIVE terminal with another wire lead.
4. Place the copper strip and the spoon in a 250-mL beaker and fill the beaker two-thirds full with copper sulfate solution.
5. Add 15 mL of concentrated sulfuric acid.
6. Remove the spoon after 3–5 min and observe the copper coating on the stainless steel.

Reactions

This demonstration is an oxidation-reduction reaction with the copper ions in solution receiving electrons from the spoon electrode and forming metallic copper, which plates on the spoon.

1. At the anode:



2. At the cathode:



Solutions

1. For best results, use a saturated copper sulfate pentahydrate solution.
2. H_2SO_4 is concentrated. CAREFUL with concentrated sulfuric acid!

Teaching Tips

NOTES

1. You can get the same effect by simply placing a metal such as zinc in a solution of copper sulfate, although the coating may appear black rather than copper-colored. In this case, the anode reaction is



2. In any electroplating reaction, the object to be plated is the cathode and the plating material is the anode.

3. You can determine the exact number of grams of copper plated out if you know the current and the time. Refer to any college-level chemistry text for this equation.
4. Oxidation occurs at the anode; reduction occurs at the cathode.
5. Try electroplating other items: paper clips, spatulas, knife blades, zinc strips, and so forth.

QUESTIONS FOR STUDENTS

1. What is oxidized and what is reduced in this demonstration?
2. Are there any practical uses for this process?
3. What other metals could be plated with copper?

104. Electrolysis of Potassium Iodide

Two simple pencil "lead" electrodes are placed in a Petri dish containing a solution of potassium iodide. When the electrodes are connected to a cell, iodine is produced and deposited on the anode.

Procedure

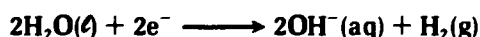
1. Place a Petri dish on a plastic sheet on an overhead projector.
2. Place a match-head-size amount of solid potassium iodide, KI, in the dish.
3. Add water to half fill the Petri dish and stir to dissolve the KI.
4. Prepare two carbon electrodes by carefully stripping a pencil to remove the pencil "lead". Each piece should be about 2-3 cm long.
5. Using alligator clips, connect a short piece of insulated wire to each electrode.
6. Attach the other ends to the poles of a 6-V cell or a 9-V battery.
7. Arrange the two electrodes so that only the carbon is in contact with the solution in the dish.
8. Add 1 drop of phenolphthalein.
9. Observe the two electrodes for evidence of reaction.

Reactions

1. At the anode:



2. At the cathode:



Phenolphthalein turns pink as the hydroxide concentration increases.

Teaching Tips

NOTES

1. Iodide is oxidized at the anode (+); water is reduced at the cathode (-).
2. You can use mechanical pencil leads.
3. Although the solubility of iodine in water is low (0.03 g/100 mL), a small amount of iodine will dissolve to produce a light brown color at the anode.
4. Hydrogen gas is produced at the cathode. The bubbles project well on the overhead projector.

QUESTIONS FOR STUDENTS

1. Why is iodine produced from potassium iodide and not potassium?
2. Why does the solution become pink in color?
3. Which electrode is the anode and which the cathode?
4. What is happening to each electrode?

105. The Aluminum-Copper Trade-Off

A tall cylinder of green-blue copper chloride solution and a strip of aluminum react to produce changes in colors, temperature, and formation of a solid.

Procedure

1. Mix 4–5 tsp of copper chloride with 500 mL of water in a tall graduated cylinder.
2. Cut a strip of aluminum from a pie pan. Twist into an interesting shape.
3. Place the aluminum in the copper chloride solution.
4. Observe and identify as many changes as possible.

Reaction

This demonstration is a replacement reaction in which the more active metal, aluminum, replaces the copper ions in solution.



Materials

1. Copper(II) chloride, CuCl_2 , crystals.
2. Aluminum strip. (Foil is too light and will disintegrate during the reaction.)

Teaching Tips

NOTES

1. During this reaction, three observations of a chemical reaction may be made: the copper chloride solution changes color from green-blue to almost colorless; the reaction evolves heat; a new colored solid forms.
2. Encourage students to write their observations. Help them avoid saying that the aluminum "is turning into a new substance" or that it is rusting!
3. The spongy copper that forms on the surface of the aluminum can be identified because it will not react with hydrochloric acid.
4. The spongy copper can be fused by a hot flame to form the more recognizable metallic copper. Take the copper to a welding facility.

QUESTIONS FOR STUDENTS

1. Identify three changes to indicate that this is a chemical reaction.
2. What change has taken place in the copper atoms that reacted?
3. What change has taken place in the aluminum atoms?
4. Will any other metal react in a similar way when placed in a solution of copper chloride?

106. Water Electrolysis in Yellow, Green, and Blue

When bromothymol blue indicator is added to water in an electrolytic apparatus, three colored regions are seen.

Procedure

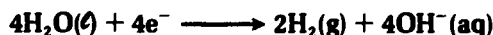
1. Prepare the needed volume of water for electrolysis by adding 1 tsp of solid sodium sulfate and a squirt of bromothymol blue indicator solution.
2. Assemble the apparatus and connect a 6-9-V dc power source.
3. Observe color changes during electrolysis and explain the sources of acid and base that form.

Reactions

1. When water is oxidized at the anode, oxygen gas and hydrogen ions form. The indicator is yellow.



2. When water is reduced at the cathode, hydrogen gas and hydroxide ions form. The indicator is blue.



3. When the hydrogen and hydroxide ions react, the solution is neutral. The indicator is green.

Materials

1. Sodium sulfate, Na_2SO_4 .
2. Bromothymol blue indicator.
3. Hoffman apparatus or U-tube fitted with two-hole stoppers and platinum or Nichrome electrodes. Bent wire in inverted test tubes will also work.

Teaching Tips

NOTES

1. The U-tube is preferred for seeing the colored regions; however, the collection of the gases is difficult.
2. A Hoffman apparatus is suited for the collection of the gases so that volumes can be compared. It is worthwhile for students to see both.
3. A 9-V calculator battery is handy for this reaction and produces the gases at a good rate.
4. You may use any salt in the water that does not react. Sodium chloride, for instance, produces chlorine gas at the anode.
5. Electrolysis of water using cheap or wasted electricity can produce hydrogen gas to be used as a fuel.

QUESTIONS FOR STUDENTS

1. What reaction takes place at the cathode? What reaction takes place at the anode?
2. Which direction do the electrons flow in the circuit?
3. How many electrons are produced for each molecule of oxygen formed?
4. How many hydrogen molecules are formed for every oxygen molecule?

107. Visible Oxidation-Reduction in Electrochemical Cells

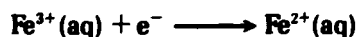
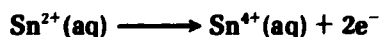
A coil of wire is stretched between two solutions in beakers. A blue color suddenly appears around the coil in one of the solutions.

Procedure

1. In beaker A, place 200 mL of iron(III) chloride, FeCl_3 , solution plus 1-2 mL of potassium hexacyanoferrate(III), $\text{K}_3\text{Fe}(\text{CN})_6$ solution.
2. In beaker B, place 200 mL of tin(II) chloride, SnCl_2 , solution.
3. Place a *salt bridge* connecting the two beakers. You may use a piece of string or yarn or twisted paper towel soaked in 1 M KCl solution.
4. Coil about 30 cm of Nichrome wire around a glass rod.
5. Immerse a glass vial in beaker A so that the opening is under the surface.
6. Stretch the wire coil in the center and place the coiled ends in the solutions, one in the vial and the other in beaker B.
7. Look for changes.
8. Repeat the procedure using a solution of iron sulfate, FeSO_4 , and 1-2 mL of potassium thiocyanate, KSCN, solution in beaker A and potassium permanganate, KMnO_4 , solution in beaker B.
9. To help explain the changes prepare two test tubes, one with iron(III) chloride solution and the other with iron(II) sulfate solution. Add a few drops of $\text{K}_3\text{Fe}(\text{CN})_6$ solution to the iron(II) sulfate and KSCN solution to the iron(III) chloride.

Reactions

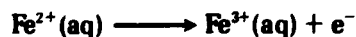
1. In the first procedure, iron(III) is reduced and tin(II) is oxidized.



The iron(II) ion reacts with $\text{Fe}(\text{CN})_6^{3-}$ to form a blue complex:



2. In the second procedure, iron(II) is oxidized and Mn(VII) is reduced.



The iron(III) ion that is formed reacts with the KSCN solution to form a red complex, FeSCN^{2+} .

Solutions

Because the concentrations are not critical, you may vary the following suggestions:

1. The iron(III) chloride concentration is 0.05 M: 4 g of FeCl_3 in 500 mL of solution.
2. Tin(II) chloride is saturated in 4–5 M HCl: Use 35 g of SnCl_2 in 500 mL of acid.
3. The iron(II) sulfate concentration is 0.1 M: 7.5 g of FeSO_4 in 500 mL of solution.
4. The potassium permanganate concentration is a 0.05 M: 4 g of KMnO_4 in 500 mL of solution.
5. Potassium ferricyanide [hexacyanoferrate(III)] is a dilute solution: Add about 1 g of $\text{K}_3\text{Fe}(\text{CN})_6$ to 100 mL of water.
6. Potassium thiocyanate is a dilute solution: Add about 0.1 g of KSCN to 100 mL of water.
7. Potassium chloride is a 1.0 M solution: 7.4 g of KCl in 100 mL of solution.

Teaching Tips

NOTES

1. The colored complexes formed in the vials will be contained in the vials.
2. Challenge students to identify the direction of electron flow in each case using the “standards” prepared in step 9 under Procedure as a guide.
3. The positive tests for iron(II) and for iron(III) in this demonstration are those commonly used for these ions.
4. Try presenting this demonstration in other convenient ways—long coils in large test tubes with a string salt bridge and no beakers or in Petri dishes on the overhead projector.

QUESTIONS FOR STUDENTS

1. What reactions caused the blue color in beaker A?
2. What reactions caused the red color in beaker B?
3. Identify the direction of the electron flow in each cell.
4. Look up the standard electrode potentials for these reactions and explain how you could predict that these reactions are spontaneous.

108. Metal Displacement: Copper and Mercury

A few drops of metallic mercury are added to a solution of copper(II) chloride in a beaker. No obvious reaction occurs. A strip of copper is added to a second beaker containing mercury(II) chloride solution, and the copper is immediately coated with metallic mercury.

Procedure

1. Fill a 250-mL beaker two-thirds full of copper(II) chloride solution.
2. Fill a second beaker two-thirds full of mercury(II) chloride solution. Clearly label the beakers.
3. Carefully place a few drops of mercury in the beaker containing the copper chloride solution. Observe for a few minutes to be sure that no reaction occurred.
4. Place a clean strip of copper in the beaker containing mercury(II) chloride solution.
5. Observe the reactions. First, note the coating of mercury on the copper strip. Then, notice that the solution becomes light green in color.

Reactions

1. Mercury will not displace copper in copper(II) chloride solution.
2. Copper will, however, displace mercury in mercury(II) chloride solution.



Solutions and Materials

1. The concentrations of the metal salt solutions is not critical. 0.1 M solutions work well.
 - a. The copper(II) chloride concentration is 0.1 M: Dissolve 13.4 g of CuCl_2 in 1 L of solution.
 - b. The mercury(II) chloride concentration is 0.1 M: Dissolve 27.1 g of HgCl_2 in 1 L of solution.
2. Metallic mercury and copper should be clean. Clean the copper strip by brushing it with a Brillo pad or dip it quickly into a dilute solution of nitric acid.

Teaching Tips

NOTES

1. This demonstration works well in Petri dishes on an overhead projector.
2. This demonstration is an example of a single-replacement reaction. Copper is a more reactive metal than mercury and would appear above it in the activity series.
3. Point out that oxidation numbers change, so this demonstration is also an example of an oxidation-reduction reaction. In double replacement (metathesis reactions), oxidation numbers may not change.
4. See Appendix 4 for proper methods of disposal of the solutions.
5. Use metallic mercury with care. It is toxic.

QUESTIONS FOR STUDENTS

1. Explain why one reaction "works", but the other does not.
2. Show how this demonstration is an oxidation-reduction reaction.
3. In what order would these two metals appear in the activity series? Why?
4. What evidence is there for a chemical reaction?

4. Oil paintings can be cleaned with hydrogen peroxide. H_2O_2 oxidizes the black lead sulfide to white lead sulfate.
5. You can easily titrate permanganate with hydrogen peroxide and determine its concentration.
6. Generally, when hydrogen peroxide acts as a reducing agent, water and oxygen are products. When it acts as an oxidizing agent, oxygen is not a product.

QUESTIONS FOR STUDENTS

1. Show, by use of equations, how hydrogen peroxide can be a reducing agent and an oxidizing agent.
2. What is the reduction product of hydrogen peroxide?
3. Why must the demonstration be done in an acid solution?
4. Explain what was reduced and what was oxidized in each reaction.

110. Manganese(III) as an Oxidizing Agent

A solution of Mn^{3+} is prepared. When added to a solution of Cr^{3+} , slow oxidation produces the colorful dichromate ion. When added to a Fe^{2+} ion, the red Mn^{3+} more rapidly oxidizes the iron to form the blue Fe^{3+} complex. Mn^{3+} is unstable in solution and is thus not commonly encountered in the laboratory.

Procedure

1. Slow oxidation:

- Place 100 mL of potassium chromium sulfate (chrome alum) solution in a 250-mL beaker. Notice the blue color of the solution.
- Add a few drops of 6 M hydrochloric acid and heat the solution until it boils.
- While stirring, add Mn^{3+} , prepared as described under Solutions, until the solution begins to change to green. Observe as the color of the dichromate is finally produced.

2. Rapid oxidation:

- Place 100 mL of iron(II) sulfate solution in a beaker.
 - Add 15–20 drops of phosphoric acid and the same amount of sodium diphenylamine sulfonate.
3. While stirring, slowly add Mn^{3+} solution. Notice the sharp color change from red to blue-violet.

Reactions

1. Slow oxidation:



2. Fast oxidation:



Solutions

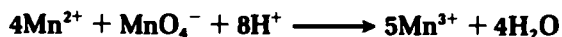
1. Make the Mn^{3+} solution as follows:

- Prepare a 0.05 M solution of manganese(II) sulfate in 3 M sulfuric acid. Add 0.8 g of MnSO_4 to 100 mL of 3 M H_2SO_4 .
 - Slowly add 0.02 M potassium permanganate solution (0.32 g/100 mL) to 60 mL of the acid solution from part a. Stir the solution.
 - Store this solution in a dark bottle in a cool place. It should be used within a few days.
- The potassium chromium sulfate (chrome alum) solution concentration is 0.02 M: Dissolve 10 g of $\text{K}[\text{Cr}(\text{SO}_4)_2] \cdot 12\text{H}_2\text{O}$ in 1 L of solution.
 - The iron(II) sulfate solution concentration is 0.02 M: Dissolve 3.34 g of FeSO_4 in 1 L of solution.
 - Sodium diphenylamine-4-sulfonate is an oxidation-reduction indicator that changes from colorless to violet upon oxidation. Ferroin may also be used as an indicator.

Teaching Tips

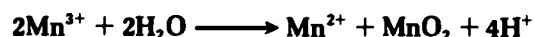
NOTES

1. Cr^{3+} is in the form $[\text{Cr}(\text{H}_2\text{O})_6]^{3+}$.
2. The reaction for preparation of the Mn^{3+} solution is



Notice that Mn^{3+} ions are produced by both oxidation of Mn^{2+} and reduction of MnO_4^- .

3. Upon heating or exposure to light, the unstable Mn^{3+} solution will disproportionate.



4. Point out that the fast oxidation involves one-electron transfer.

QUESTIONS FOR STUDENTS

1. Show by reaction that Mn^{3+} acts as an oxidizing agent.
2. Why is the Mn^{3+} ion unstable?
3. What seems to influence the rate of an oxidation-reduction reaction?
4. Why is the Mn^{3+} solution in an acid solution?

111. Indigo: The Oldest Dye

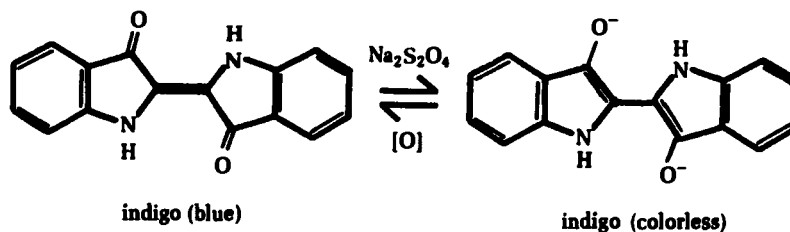
A piece of white cotton is dipped into a light yellow solution. When the cloth is removed from the solution, it becomes deep blue within a few seconds.

Procedure

1. Prepare solution A as described under Solution and Materials.
2. Add a match-head amount of sodium hydrosulfite to 300 mL of water at 50 °C. Allow the solution cool for a few minutes.
3. Add the contents of the test tube from step 1 to the water. The solution should be clear and perhaps light yellow.
4. Dip a white cotton handkerchief, or other cloth, into the solution.
5. Notice that the cotton cloth remains colorless.
6. Using forceps, remove the cloth from the solution. Notice that it turns deep blue after a few seconds.

Reactions

Indigo is blue (and insoluble in water) in its usual form. For this demonstration, the indigo is first reduced to its colorless form with sodium hydrosulfite.



The colorless reduced form is oxidized in air to produce the blue form of the dye.

Solutions and Materials

1. Solution A:

- a. Prepare a concentrated sodium hydroxide solution by dissolving 1.5 g of NaOH in 5 mL of water in a test tube.
 - b. Add 1.5 g of solid indigo to the test tube.
 - c. While shaking the test tube, add 20 mL of water.
 - d. Place the test tube in a water bath at 50 °C.
 - e. When the mixture in the test tube reaches this temperature, slowly add 2.0 g of sodium hydrosulfite, shaking the test tube with each addition.
 - f. With constant stirring, keep the test tube at this temperature for 20-30 min. The solution should now be colorless or light yellow, and all the indigo should be dissolved.
2. Indigo (indigotin) is a complex molecule with the chemical formula $C_{16}H_{10}N_2O_2$.
 3. Sodium hydrosulfite, $Na_2S_2O_4$, is also called sodium dithionite.

Teaching Tips**NOTES**

1. This reaction has historical implications, because it is one of the oldest known dyes. Mummy cloth, over 5,000 years old contain this dye.
2. In the ancient world, the dye was "fermented" in large vats to reduce it to the colorless form. Now, we use sodium hydrosulfite for this purpose.
3. The insoluble blue pigment is "locked" within the fabric when it is oxidized by oxygen in air.
4. The colorless form of indigo is sometimes called the *leuco* form.

QUESTIONS FOR STUDENTS

1. Show how this demonstration is an oxidation-reduction reaction. What is oxidized and what is reduced?
2. What properties of indigo make it act as a "dye"?
3. Indigo is referred to as a "vat" dye. Can you suggest why?

112. Colorful Oxidation of Alcohols

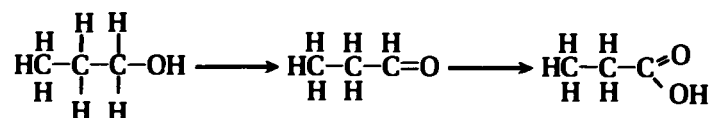
When a yellow solution is added to each of three alcohols, curiously, the solution becomes green and then blue in only two of the alcohols.

Procedure

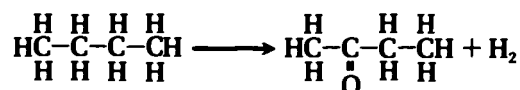
1. Place 10 mL of propyl alcohol in a large test tube, 10 mL of *sec*-butyl alcohol in another, and 10 mL of *tert*-butyl alcohol in a third test tube.
2. With student helpers, add 15 mL of acidified potassium dichromate solution to each test tube at the same time.
3. Stir each solution quickly.
4. Observe color changes and compare the test tubes.

Reactions

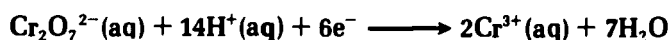
1. The primary alcohol, propyl alcohol, is oxidized to the aldehyde, propanal, and then to the acid, propanoic acid.



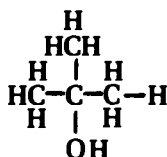
2. The secondary alcohol, *sec*-butyl alcohol, is oxidized to the ketone, 2-butanone.



3. In reactions 1 and 2, the dichromate ion is reduced to chromium(III), which appears green in the mixture and is blue or violet when hydrated.



4. The tertiary alcohol, *tert*-butyl alcohol, cannot be oxidized because no hydrogen atoms are on the carbon to which the OH group is bonded.



Materials and Solutions

1. Propyl alcohol (1-propanol).
2. *sec*-Butyl alcohol (2-butanol).
3. *tert*-Butyl alcohol (2-methyl-2-propanol).
4. Acidified dichromate: 0.015 M potassium dichromate, $K_2Cr_2O_7$, in 3 M sulfuric acid, H_2SO_4 . Add 2.2 g of potassium dichromate to 3 M sulfuric acid to make 500 mL of solution.

Teaching Tips

NOTES

1. The complete color change takes about 60 s depending upon the room temperature.
2. Butyl alcohol can be substituted for the propyl alcohol; however, butyl alcohol has a low solubility in water and the reaction will be slower.
3. The position of the OH group is found to be either primary or secondary by this test. You might try some other alcohols such as ethanol and isopropyl alcohol.
4. Try an additional test for structure by adding several drops of 12 M HCl to a sample of each alcohol. This time only the *tert*-butyl alcohol reacts and produces a milky emulsion.

QUESTIONS FOR STUDENTS

1. Describe the reactions you observed.
2. In what way are the two alcohols that reacted with the dichromate similar to each other and different from the alcohol that did not react?
3. What gas did you see being produced during the reaction?
4. Use structural formulas to show where the atoms producing the gas came from.

Appendix 1. Cross-Reference of Demonstrations and Chemical Topics

Topic	Demonstration
Historical note	1, 2, 48, 62, 111
Environmental and applied chemistry	22, 29, 34, 43, 80, 87, 90, 91, 94, 98, 99, 102, 103, 111
Air pressure	1, 14
Density	6, 7, 11, 18, 22, 35, 49
Phase change	8, 10, 13, 16, 22, 28, 37, 46, 64
Molecular mass	15, 21, 22, 38, 103
Properties of water	7, 9, 48
Decomposition reaction	4, 20, 23, 24, 58, 80
Composition reaction	24, 30, 32, 58, 79
Gases	2, 11, 12, 13, 15, 17, 18, 19, 20, 21, 22, 23, 24, 25, 26, 27, 29, 30, 35, 75, 87, 94
Reactions of elements	4 (Cu), 17 (Al), 24 (O ₂), 27 (Ca), 28 (S), 29 (Al), 30 (Cl, Na), 31 (Mg), 32 (Ca, S), 33 (Br, I), 34 (Fe), 35 (Cu, Zn), 36 (Cu), 37 (I ₂), 54 (Fe), 79 (Cu, S), 87 (S), 99 (Fe), 100 (Hg), 101 (Fe), 103 (Cu), 105 (Al, Cu), 108 (Cu, Hg)
Complex ions	36, 38, 39, 40, 41, 42, 44, 67, 88, 95
Energy change	3, 5, 19, 22, 24, 30, 31, 32, 38, 53, 54, 55, 56, 57, 58, 59, 60, 61, 64
Solubility: precipitates	4, 8, 43, 44, 62, 63, 65, 66, 67, 68, 69, 70, 73, 94
Solubility: polar-nonpolar	3, 33, 47, 49, 50, 52, 53, 95
Crystals	28, 29, 37, 46, 64, 68, 73
Reaction mechanisms	23, 26, 61, 74, 76, 77, 78, 81, 82, 84, 86, 110
Catalysts	23, 56, 79, 80, 81, 82, 83
Equilibrium	42, 44, 46, 66, 85, 88, 91, 94, 95
Acids and bases	2, 4, 17, 18, 20, 21, 25, 27, 29, 35, 55, 70, 72, 75, 76, 85, 86, 87, 89, 90, 91, 92, 93, 94, 95, 106
pH	86, 90, 91, 93
Oxidation-reduction	2, 4, 5, 24, 27, 29, 30, 31, 32, 33, 36, 45, 51, 54, 59, 74, 82, 83, 95, 96, 97, 98, 99, 100, 101, 102, 103, 104, 105, 106, 107, 108, 109, 110, 111, 112
Organic reactions	3, 51, 52, 53, 57, 58, 60, 77, 81, 84, 111, 112

Appendix 2. Properties and Preparation of Laboratory Acids and Bases

Parameter	Ammonium Hydroxide (NH_4OH)	Acetic Acid ($\text{HC}_2\text{H}_3\text{O}_2$)	Hydrochloric Acid (HCl)	Nitric Acid (HNO_3)	Sulfuric Acid (H_2SO_4)
Dilute this volume (in milliliters) of concentrated reagent to 1 L to make a 1.0 M solution	67.5	57.5	83.0	64.0	56.0
Dilute this volume (in milliliters) of concentrated reagent to 1 L to make a 3.0 M solution	200	172	249	183	168
Dilute this volume (in milliliters) of concentrated reagent to 1 L to make a 6 M solution	405	345	496	382	336
Normality of concentrated reagent	14.8	17.4	12.1	15.7	36.0
Molecular weight	35.05	60.05	36.46	63.02	98.08
Specific gravity of concentrated reagent	0.90	1.05	1.19	1.42	1.84
Approximate percentage in concentrated reagent	57.6	99.5	37.0	69.5	96.0

Note: To make normal solutions, use the same amount of reagent shown. However, to make a normal solution of sulfuric acid, use half the amount of reagent indicated. Example: dilute 28.0 mL of concentrated sulfuric acid to make 1 L of 1.0 N sulfuric acid solution.

Appendix 3. Equipment and Reagent List

This list includes the reagents and the equipment that are needed for each demonstration. Balances for weighing have not been listed. This list is for your convenience. You should always look at the individual demonstration and determine the concentration that works best for you on the basis of the purity and age of your stock reagents.

The numbers in italics indicate the number of the demonstration.

Equipment

Alligator clips	59, 101, 103, 104	Gas collecting tube, 50 mL	6
Aluminum foil	17, 29	Gas delivery tube	19
Aluminum pie pan	105	Glass bend	24, 98
Aluminum soft drink can	14, 29	Glass plate	4, 24, 30
		Glass tubing	15, 37
Baggie	13, 54	Glass vial	107
Balloon	10, 12, 17, 75	Glass wool	21
Banana	10	Graduated cylinder, 10 mL	18, 26, 33, 36, 53, 61, 80, 81, 82, 84, 86, 93, 95, 97, 101, 111, 112
Battery, 1.5 V	101, 103	Graduated cylinder, 50 mL	4, 5, 8, 10, 17, 23, 25, 29, 30, 53, 57, 66, 77, 82, 84
Battery, 6 V	104	Graduated cylinder, 100 mL	16, 39, 40, 41, 43, 50, 55, 57, 66, 69, 70, 75, 77, 79, 82, 85, 87, 89, 90, 91, 93, 94, 96, 105, 107, 109, 110
Battery, 9 V	106	Graduated cylinder, 250 mL	88
Battery jar, 2.0 L	1, 7	Graduated cylinder, 500 mL	25
Beaker, 100 mL	8, 35, 45, 62, 86, 92, 97	Graduated cylinder, 1000 mL	7, 18, 88
Beaker, 250 mL	4, 5, 7, 16, 25, 27, 28, 29, 31, 34, 37, 38, 40, 41, 43, 51, 53, 57, 59, 67, 69, 71, 72, 78, 82, 84, 85, 89, 91, 93, 95, 96, 103, 107, 108, 109, 110, 111	Grapes	10
Beaker, 400 mL	70		
Beaker, 500 mL	9, 12, 15, 76, 77, 79, 90, 101, 111	Handkerchief	42, 43, 111
Beaker, 600 mL	74, 95, 105	Hollow rubber ball	10
Buret	48, 70	Hot plate	73
Can of diet soda	7	Ice	7, 8, 12, 13, 29, 55, 57, 78
Can of regular soda	7	Index card	50
Candle	17, 58	Iron-fortified cereal	34
Carnation	10	Iron nail	10, 36, 99, 101
Chalk	60		
Conductivity apparatus	70, 71	Large opaque bottle	23
Cotton balls	56, 57	Lighter fluid	3
		Litmus paper	21, 73
Deflagrating spoon	30, 87		
Dewar flask, wide	10	Magnetic stirrer	34, 76, 77, 90
Dishwashing detergent	25	Masking tape	13
		Meter stick	17
Egg	7, 72	Milk of Magnesia tablet	90
Electrolysis setup	106	Mothballs	18
Erlenmeyer flask, 250 mL	10, 17, 50, 66, 75		
Erlenmeyer flask, 500 mL	2, 3, 24	Overhead projector	46, 47, 63, 104
Erlenmeyer flask, 1000 mL	87		
Erlenmeyer flask and stopper, 500 mL	68	Pencil (2)	104
Evaporating dish	8	Penny	2
		Penny (made since 1983)	35
Filter paper	29, 73	Petri dish	47, 52, 63, 99, 100, 104
Florence flask, 500 mL	13, 30	Pie pan	10
Florence flask and stopper, 250 mL	64	Plastic rod	48
Florence flask and stopper (2), 1000 mL	61	Plastic soft drink bottle	11, 14, 49
Food color	1, 7, 15, 49, 53	Plastic wash bottle	11, 14, 49
		Plastic wrap	72

Porous clay cup and stopper 15
 Potato 80, 81
 Pyrex test tube, 16 mm \times 150 mm 37
 Pyrex test tube and stopper, one hole 24

Rubber tubing 24, 30, 80, 88

Sand 8, 22, 102
 Sewing needle 9
 Sidearm flask and stopper 30
 Soap solution 9
 Spaghetti 18
 Stainless steel spoon 103
 Steel wool 24
 Steel wool, 000 97
 Stirring rod, long 88
 Stopper, one hole 37, 98
 String 107
 Styrofoam-brand cup 52, 55
 Syrup, clear 72

Tea bag, empty 23
 Test tube 19, 21, 27, 28, 32, 36, 44, 65, 74, 80,
 81, 83, 97, 98, 102, 111, 112
 Test tube, 25 mm \times 200 mm 26
 Test tube and stopper 33, 60
 Test tube and stopper, one hole 19, 80
 Thermometer 5, 55, 73, 74, 75, 78, 79,
 80, 111
 Thermometer, to 200 °C 16, 37
 Tissue paper 9
 Tongs 31
 Triangular file 35

U-tube 71
 Ultraviolet light 60

Watch glass 38, 45, 46, 57
 Water trough 19, 24, 80
 White writing surface 60
 Wooden splint 27
 Wool 48

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Acetic acid, 1.0 M 85
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 Acetone 6, 52, 84
 Alcohol 16
 Aluminum wire, 18 gauge 4
 Ammonia 41
 Ammonia, concentrated 44, 73, 95
 Ammonium chloride 20, 21, 71
 Ammonium sulfate 26
 Anthracene 60

Barium chloride 65
 Barium hydroxide 70
 Benzaldehyde 84
 Bismuth, metal cube 16

Bleach, commercial 33
 Bromothymol blue 86, 88, 106
 2-Butyl alcohol 112
 Calcium chloride 62, 67
 Calcium hydroxide 20, 43, 94, 98
 Calcium hypochlorite 30
 Calcium, metal 27, 32
 Carbon, powdered 98
 Catechol 81
 Chromium(III) oxide 45
 Cinnamaldehyde 84
 Cobalt(II) chloride 42
 Cobalt nitrate 41
 Copper chloride 65, 105, 108
 Copper nitrate 41
 Copper oxide 98
 Copper sulfate 36, 38, 39, 47, 63, 79, 80, 95,
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Copper wire, fine 4
 Copper wire, heavy 59
 Copper, powdered 79
 Copper, strip 103, 108
 Cottonseed oil 37
 Cresol red 86
 o-Cresolphthalein 86
 Cyclohexane 51
 1,4-Cyclohexanediol 81
 Cyclohexene 51

Diaminoethane 40
 Dichloromethane 61
 Dimethylgloxime 67
 Dry ice 31, 71, 94

Ethanol 7, 38, 92

Formaldehyde 76, 86

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 Glycerol 16
 Glycerine 46

Hexane 48
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 Hydrochloric acid, 1.0 M 33, 89, 90, 93
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 Hydrochloric acid, 3.0 M 17, 75
 Hydrochloric acid, 4.8 M 88
 Hydrochloric acid, 6.0 M 4, 15, 30, 35, 59
 Hydrochloric acid, concentrated 18, 39,
 95, 110
 Hydrogen peroxide, 3% 78, 80, 109
 Hydrogen peroxide, 30% 23, 61

Indigo 111
 Iodine 8, 37, 47
 Iron(II) chloride 95, 107
 Iron(II) sulfate 67, 107, 110
 Iron, powdered 54
 Isoviolanthrone 61

Lead iodide	68	Silver nitrate	44, 66, 83, 97
Lead nitrate	43, 65	Sodium	30
Magnesium ribbon	31, 59	Sodium acetate	19, 85
Magnesium sulfate	62	Sodium acetate trihydrate	46, 64
Magnesium, powdered	102	Sodium bicarbonate	18, 25
Manganese dioxide	24, 45	Sodium bisulfite	5, 76
Manganese(II) sulfate	83, 110	Sodium bromate	77
Mercury	100, 108	Sodium bromide	77
Mercury(II) chloride	65, 108	Sodium carbonate	63, 65, 75
Methyl orange	85, 88, 91, 93	Sodium chloride	7, 17, 36, 54, 66
2-Methyl-2-propyl alcohol	112	Sodium chromate	66
Methyl red	77	Sodium dichromate	5
Methyl violet	93	Sodium diphenylaminesulfonate	110
Nichrome wire	107	Sodium fluoride	95
Nickel chloride	40, 65	Sodium hydrogen sulfite	96
Nickel nitrate	41	Sodium hydrosulfite	111
Nickel sulfate	67	Sodium hydroxide	19, 45, 62, 85, 88, 89, 91, 92, 111
Nitric acid, 1.0 M	100	Sodium metabisulfite	86
Nitric acid, 6.0 M	4, 44, 95	Sodium perchlorate	26
Nitric acid, concentrated	2, 57, 97, 103	Sodium peroxide, fresh	56
Nitric acid, dilute	45, 59	Sodium silicate	69
Nitrocellulose	57	Sodium sulfate	65, 106
Nitrogen, liquid	10	Sodium sulfite	76, 86
Oxalic acid	73, 74, 82	Sodium tetraborate	53
Oxalyl chloride	61	Sodium thiosulfate	78
Perylene	61	Sodium thiosulfate pentahydrate	46
Phenol	77, 81	Starch, soluble	78
Phenolphthalein	27, 69, 70, 76, 91, 95, 99, 101, 104	Strontium chloride	69
Phenyl salicylate	46	Sucrose	58
Phosphoric acid, concentrated	110	Sulfur	28, 32, 87
Poly(vinyl alcohol)	53	Sulfur, powdered	79
Potassium acid phosphate	91	Sulfuric acid, 0.2 M	77
Potassium bromate	33	Sulfuric acid, 1.0 M	70
Potassium bromide	33	Sulfuric acid, 3.0 M	63, 110
Potassium carbonate	62, 67	Sulfuric acid, 9.0 M	29, 55
Potassium chlorate	24, 58	Sulfuric acid, concentrated	57, 58, 82, 103, 109
Potassium chloride	33, 107	Tetracene	61
Potassium chromate	8	Thymol blue	86
Potassium chromium sulfate	10	Thymolphthalein	86, 92, 95
Potassium dichromate	43, 96, 100, 112	Tin, metal cube	16
Potassium ferricyanide	67, 99, 101, 107	Tin(II) chloride	97
Potassium hydroxide	29, 62, 84, 91	Toluene	60
Potassium iodate	33	Trichlorotrifluoroethane	33, 47, 50
Potassium iodide	23, 33, 65, 78, 95, 104, 109	Universal indicator	87, 90
Potassium nitrate	26, 45	Vermiculite	54
Potassium perchlorate	26	Violanthrone	61
Potassium permanganate	51, 74, 82, 107, 109, 110	Zinc, dust	22
Potassium persulfate	83	Zinc, metal	15
Potassium thiocyanate	95, 107	Zinc, mossy	18
1-Propyl alcohol	112	Zinc chloride	89
Rubrene	61		

Appendix 4. Safety and Disposal

The following information is meant to be a general guide for you to consider before performing a demonstration and disposing of the chemical compounds formed or used in each demonstration.

ALWAYS wear safety goggles when you are performing a demonstration, and insist that any students assisting you wear them also. If the demonstration requires a hood or adequate ventilation, it should not be performed otherwise. Take special care when making solutions and diluting acids. Remember, always add acid to water when making dilutions. Do not exceed the amounts of chemicals recommended for a demonstration and perform each demonstration only under the conditions specified in the procedure.

Check with your district science supervisor or state department of education to see which substances are on the "banned" list for your schools and if specific directions for disposal of chemical wastes are available. The specifications and requirements vary from state to state. Generally, you can dispose of waste acids and bases as follows:

1. Neutralize acids with ammonium hydroxide or a sodium bicarbonate solution and flush down the drain with copious amounts of water.
2. Neutralize bases with a citric acid solution or dilute hydrochloric acid and flush down the drain with copious amounts of water.

The following disposal procedures for chemicals used in these demonstrations are suggested in *Prudent Practices for Disposal of Chemicals from Laboratories* (National Academy: Washington, DC, 1983).

Inorganic Chemicals

Chemical	Disposal Procedure
Barium	Precipitate as the sulfate or carbonate and dispose of in an EPA-regulated landfill.
Cadmium	Add sodium sulfide solution to the metal ion solution to precipitate the heavy metal sulfide. Make this solution basic by adding ammonium hydroxide. Wash and dry the precipitated metal sulfide. Dispose of the precipitate in an EPA-regulated landfill. Flush the remaining solution down the drain.
Cations (other than those listed)	If in small quantities, flush down the drain with excess water.
Chromate (CrO_4^{2-} , $\text{Cr}_2\text{O}_7^{2-}$)	Check for regulations in effect in your state. Usually, you can dispose of chromates by adjusting the pH to 3 with sulfuric acid, adding 50% excess sodium bisulfite until a temperature increase results, and precipitating with NaOH and disposing of the solid in an EPA-regulated landfill.
Chromium(III)	Precipitate as the hydroxide or carbonate and dispose of in an EPA-regulated landfill.
Cobalt	Same as for Cadmium.
Cyanides	This must be done in a hood. Add 50% excess bleach at a temperature of 0-10 °C. When heat is no longer evolved, allow the mixture to stand at room temperature for several hours and then flush down the drain with excess water.
Ferricyanide [$\text{Fe}(\text{CN})_6^{3-}$]	Precipitate with Fe^{2+} . Dispose of in an EPA-regulated landfill.
Ferrocyanide	Precipitate as Fe^{3+} . Dispose of in an EPA-regulated landfill.

Lead	Same as for Cadmium.
Manganese(II)	Same as for Cadmium.
Mercury	Same as for Cadmium.
Nickel	Same as for Cadmium.
Silver	Same as for Cadmium.
Vanadium	Same as for Cadmium.

Organic Chemicals

The following organic compounds must be incinerated in an EPA-regulated incinerator and should NOT be flushed down the drain:

- Alcohols and phenols
- Aldehydes
- Amides
- Amines
- Carboxylic acids
- Esters
- Ethers
- Halogenated hydrocarbons
- Hydrocarbons
- Ketones

Appendix 5. Periodic Chart of the Elements

1	2	3	4	5	6	7	8-10				11	12	13	14	15	16	17	18
IA	IIA	IIIB	IVB	VB	VIB	VII B	VIII				IB	IIB	IIIA	IVA	VA	VIA	VIIA	NOBLE GASES
1 H 1.0079†	<div> <div>4 Be 9.01218</div> <div>12 Mg 24.305</div> </div>												5 B 10.81	6 C 12.011	7 N 14.0067	8 O 15.9994†	9 F 18.99848	10 Ne 20.179†
3 Li 6.941†													13 Al 26.98154	14 Si 28.086†	15 P 30.97376	16 S 32.06	17 Cl 35.452	18 Ar 39.948†
11 Na 22.98977							26 Fe 55.847†	27 Co 58.9332	28 Ni 58.70†	29 Cu 63.546†	30 Zn 65.38		31 Ga 69.72	32 Ge 72.59†	33 As 74.9216	34 Se 78.96†	35 Br 79.904	36 Kr 83.80
19 K 39.098†	20 Ca 40.08	21 Sc 44.9559	22 Ti 47.88†	23 V 50.9414†	24 Cr 51.996	25 Mn 54.9380	26 Fe 55.847†	27 Co 58.9332	28 Ni 58.70†	29 Cu 63.546†	30 Zn 65.38		31 Ga 69.72	32 Ge 72.59†	33 As 74.9216	34 Se 78.96†	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678†	38 Sr 87.62	39 Y 88.9058	40 Zr 91.22	41 Nb 92.9064	42 Mo 95.94†	43 Tc 98.9064	44 Ru 101.07†	45 Rh 102.9055	46 Pd 106.4	47 Ag 107.868	48 Cd 112.40		49 In 114.82	50 Sn 118.69†	51 Sb 121.75†	52 Te 127.60†	53 I 126.9045	54 Xe 131.29
55 Cs 132.9054	56 Ba 137.34†	57 *La 138.9055†	72 Hf 178.49†	73 Ta 180.9479†	74 W 183.85†	75 Re 186.207	76 Os 190.2	77 Ir 192.22†	78 Pt 195.08†	79 Au 196.9665	80 Hg 200.59†		81 Tl 204.37†	82 Pb 207.2	83 Bi 208.9804	84 Po (210)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.0254	89 ‡Ac (227)	104 Unq (281)	105 Unp (282)	106 Unh (282)	107 Uns (282)	108 Uno (285)	109 Une (286)										

* Lanthanide Series

58 Ce 140.12	59 Pr 140.9077	60 Nd 144.24†	61 Pm (147)	62 Sm 150.4	63 Eu 151.96	64 Gd 157.25†	65 Tb 158.9254	66 Dy 162.50†	67 Ho 164.9304	68 Er 167.26†	69 Tm 168.9342	70 Yb 173.04†	71 Lu 174.97
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‡ Actinide Series

90 Th 232.0381	91 Pa 231.0369	92 U 238.0289	93 Np 237.0482	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (254)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)
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Appendix 6. Atomic Weights of the Elements

Name	Symbol	Atomic Number	Atomic Weight	Name	Symbol	Atomic Number	Atomic Weight	Name	Symbol	Atomic Number	Atomic Weight
Actinium	Ac	89	227.03	Helium	He	2	4.00	Radium	Ra	88	226.02
Aluminum	Al	13	26.98	Holmium	Ho	67	164.93	Radon	Rn	86	(222)
Americium	Am	95	(243)	Hydrogen	H	1	1.01	Rhenium	Re	75	186.21
Antimony	Sb	51	121.75	Indium	In	49	114.82	Rhodium	Rh	45	102.90
Argon	Ar	18	39.95	Iodine	I	53	126.90	Rubidium	Rb	37	85.47
Arsenic	As	33	74.92	Iridium	Ir	77	192.22	Ruthenium	Ru	44	101.07
Astatine	At	85	(210)	Iron	Fe	26	55.85	Samarium	Sm	62	150.36
Barium	Ba	56	137.33	Krypton	Kr	36	83.80	Scandium	Sc	21	44.96
Berkelium	Bk	97	(247)	Lanthanum	La	57	138.90	Selenium	Se	34	78.96
Beryllium	Be	4	9.01	Lawrencium	Lr	103	(260)	Silicon	Si	14	28.08
Bismuth	Bi	83	208.98	Lead	Pb	82	207.2	Silver	Ag	47	107.87
Boron	B	5	10.81	Lithium	Li	3	6.94	Sodium	Na	11	22.99
Bromine	Br	35	79.90	Lutetium	Lu	71	174.97	Strontium	Sr	38	87.62
Cadmium	Cd	48	112.41	Magnesium	Mg	12	24.30	Sulfur	S	16	32.06
Calcium	Ca	20	40.08	Manganese	Mn	25	54.94	Tantalum	Ta	73	180.95
Californium	Cf	98	(251)	Mendelevium	Md	101	(258)	Technetium	Tc	43	(98)
Carbon	C	6	12.01	Mercury	Hg	80	200.59	Tellurium	Te	52	127.60
Cerium	Ce	58	140.12	Molybdenum	Mo	42	95.94	Terbium	Tb	65	158.92
Cesium	Cs	55	132.90	Neodymium	Nd	60	144.24	Thallium	Tl	81	204.38
Chlorine	Cl	17	35.45	Neon	Ne	10	20.18	Thorium	Th	90	232.04
Chromium	Cr	24	51.996	Neptunium	Np	93	237.05	Thulium	Tm	69	168.93
Cobalt	Co	27	58.93	Nickel	Ni	28	58.69	Tin	Sn	50	118.69
Copper	Cu	29	63.55	Niobium	Nb	41	92.91	Titanium	Ti	22	47.88
Curium	Cm	96	(247)	Nitrogen	N	7	14.01	Tungsten	W	74	183.85
Dysprosium	Dy	66	162.50	Nobelium	No	102	(259)	Unnilennium	Une	109	(266)
Einsteinium	Es	99	(252)	Osmium	Os	76	190.2	Unnilhexium	Unh	106	(262)
Erbium	Er	68	167.26	Oxygen	O	8	15.999	Unniloctium	Uno	108	(265)
Europium	Eu	63	151.96	Palladium	Pd	46	106.42	Unnilpentium	Unp	105	(262)
Fermium	Fm	100	(257)	Phosphorus	P	15	30.97	Unnilquadium	Unq	104	(261)
Fluorine	F	9	18.998	Platinum	Pt	78	195.08	Unnilseptium	Uns	107	(262)
Francium	Fr	87	(223)	Plutonium	Pu	94	(244)	Uranium	U	92	238.03
Gadolinium	Gd	64	157.25	Polonium	Po	84	(209)	Vanadium	V	23	50.94
Gallium	Ga	31	69.72	Potassium	K	19	39.1	Xenon	Xe	54	131.29
Germanium	Ge	32	72.59	Praseodymium	Pr	59	140.91	Ytterbium	Yb	70	173.04
Gold	Au	79	196.97	Promethium	Pm	61	(145)	Yttrium	Y	39	88.91
Hafnium	Hf	72	178.49	Protactinium	Pa	91	231.04	Zinc	Zn	30	65.38
								Zirconium	Zr	40	91.22

NOTE: Based on carbon-12. Numbers in parentheses are mass numbers of the most stable isotopes of radioactive elements.

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